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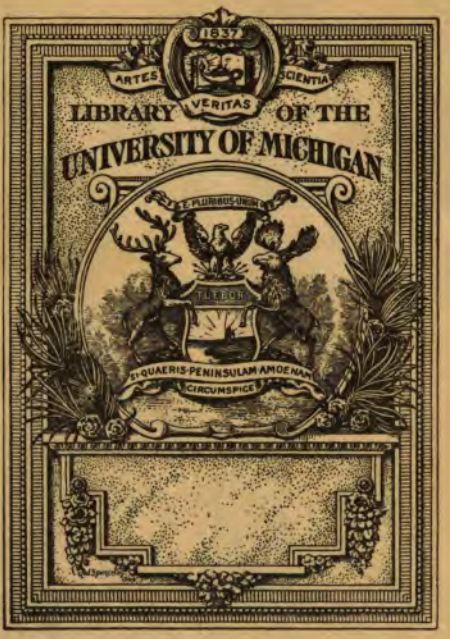
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CHEMISTRY
AND ITS RELATIONS TO DAILY LIFE

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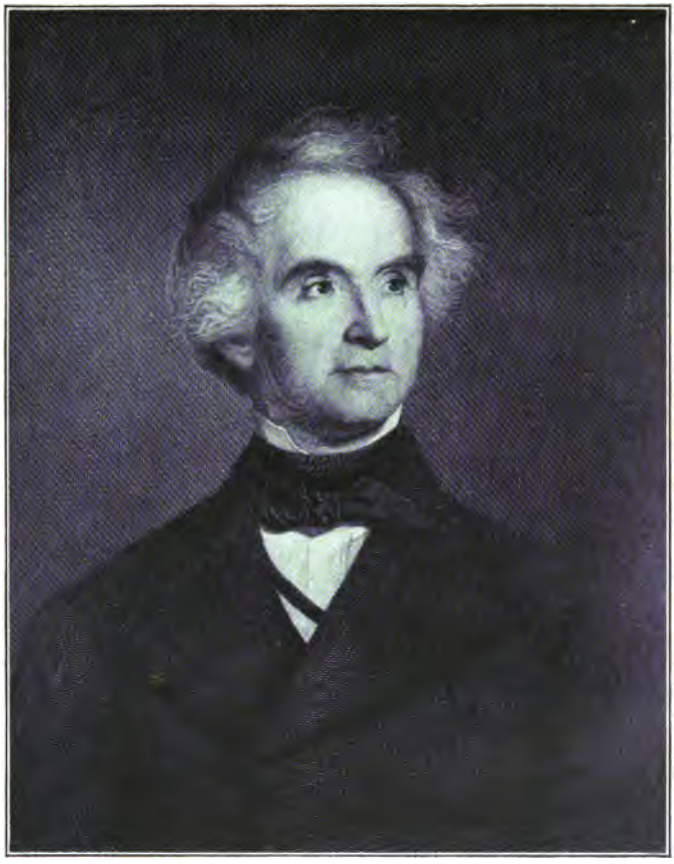

JUSTUS VON LIEBIG. 1803–1873.
The father of Agricultural Chemistry. He was the first to teach a rational basis for fertilizing the land.

CHEMISTRY

AND ITS RELATIONS TO DAILY LIFE

A TEXTBOOK FOR STUDENTS
OF AGRICULTURE AND HOME ECONOMICS
IN SECONDARY SCHOOLS

BY

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PREFACE

THIS book is intended to represent a year's work for students of agriculture and home economics in secondary schools. The aim has been to make the subject matter thoroughly practical in character and to present it in an interesting and simple way so that the student may grasp it. At the same time, the delineation has been made on sound scientific lines, and it will require patient and continuous application on the part of the student to accomplish the work intelligently. Useful facts have naturally been placed in the foreground, and no more theory has been presented than necessary. Chemical formulas have been introduced to some extent, but merely as an aid in expressing facts in simple, compact, and convenient form. Atomic and molecular theories have not been presented, for they are not calculated to aid students at this stage of advancement. While the book is not intended as a preparatory course for colleges, yet those who have completed it successfully will doubtless have gained information and sound scientific training which will compare favorably with that received in the pursuit of courses usually taken in preparatory schools.

The laboratory experiments detailed in Chapter XXI are to be performed by the student *in connection with the study of each of the preceding chapters.* The teacher should supervise this work carefully and require the student to keep neat, accurate records in a notebook. At the end of Chapter XXII are given lists of apparatus and chemicals that will be needed. Addresses of a few well-known firms from whom

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such supplies may be obtained have also been furnished for the convenience of the teacher.

The questions at the end of each chapter will help in reviewing the salient points of that chapter. The teacher will, of course, raise many additional questions, especially in connection with the laboratory exercises, upon which special stress should be laid.

While this book is primarily intended as a textbook to be used by pupils in schools under the supervision of a teacher, there are doubtless many others who will find it helpful as a volume of general information for home reading and study. Any corrections or suggestions that may be useful in preparing future editions will be gratefully received.

THE AUTHORS.

MADISON, WISCONSIN,
February 20, 1913.

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CHEMISTRY AND ITS RELATIONS TO DAILY LIFE

CHAPTER I

GENERAL FUNDAMENTAL CONSIDERATIONS

ANIMALS, plants, water, rocks, soil, and the air are the things that make up the earth and its atmosphere. These have been the subject of study since earliest times, and a considerable amount of knowledge has consequently been gathered concerning living beings, plants, and animals, and their relations to water, the air, and the solid portions of the earth's crust. All of this knowledge is the result of observation by means of our senses, sight, hearing, smell, taste, touch, and experience of changes of temperature. For many centuries the results of such observations have gradually been collected, corrected, and recorded in books and manuscripts. Such carefully gathered and systematically arranged knowledge is called **Science.** In other words, *science is systematized knowledge resulting from careful observation made with a definite purpose in view.* The opinions that men have formed as a result of reflection, that is, thinking over such observations, also form a part of science, and indeed such opinions are commonly recorded in scientific books. Now it is clear that these opinions may differ somewhat, and sometimes they differ quite widely, especially if the observations have been made with insufficient care, or if not enough observations have been made to warrant drawing a fairly safe conclusion.

The opinions then change from time to time as better observations and more of them accumulate. Consequently the opinions, that is to say, the **theories** or theoretical views, are the changeable, the ephemeral part of science, though, to be sure, the observations too may be improved and multiplied, as already stated.

The Greek philosopher Aristotle (384–322 B.C.) taught that water, air, earth, and fire are the elements of which all other things are composed, and indeed this view was held for many centuries. Even now in literature, in the descriptions of storms, conflagrations, and the like, we sometimes still speak of the " battle of the elements." Such allusions, to be sure, are now only figures of speech, for the elements of the ancient Greeks are not at all the elementary constituents of which things are composed. In fact a careful study of all of the available materials on the earth, including the atmosphere has shown that *there are about eighty substances that cannot be made from other substances and cannot themselves be resolved into anything else.* For example, iron is one of these eighty substances. No one has ever been able to make iron from other substances that do not already contain it; neither has it been possible to produce other substances from iron alone. Iron is consequently spoken of as an **element** or an elementary substance. Similarly carbon is an element. This substance we have in almost pure form in the diamond, graphite, and the better grades of charcoal.

In Table 1 given on the opposite page is a list of the chemical elements and the abbreviations or symbols by means of which they are very often designated.

In the list there are many commonly known substances. These have been designated by means of heavy type. These more common elements number only about twenty-five, and indeed the most important elements are scarcely

TABLE 1

THE CHEMICAL ELEMENTS AND THEIR SYMBOLS

ELEMENT	SYMBOL	ELEMENT	SYMBOL
Aluminum	**Al**	Neodymium	Nd
Antimony	Sb	Neon	Ne
Argon	A	**Nickel**	**Ni**
Arsenic	**As**	Niton (radium	
Barium	Ba	emanation)	Nt
Bismuth	Bi	**Nitrogen**	**N**
Boron	B	Osmium	Os
Bromine	Br	**Oxygen**	**O**
Cadmium	Cd	Palladium	Pd
Cæsium	Cs	**Phosphorus**	**P**
Calcium	**Ca**	**Platinum**	**Pt**
Carbon	**C**	**Potassium**	**K**
Cerium	Ce	Praseodymium	Pr
Chlorine	**Cl**	Radium	Ra
Chromium	Cr	Rhodium	Rh
Cobalt	Co	Rubidium	Rb
Columbium	Cb	Ruthenium	Ru
Copper	**Cu**	Samarium	Sa
Dysprosium	Dy	Scandium	Sc
Erbium	Er	Selenium	Se
Europium	Eu	**Silicon**	**Si**
Fluorine	F	**Silver**	**Ag**
Gadolinium	Gd	**Sodium**	**Na**
Gallium	Ga	Strontium	Sr
Germanium	Ge	**Sulphur**	**S**
Glucinum	Gl	Tantalum	Ta
Gold	**Au**	Tellurium	Te
Helium	He	Terbium	Tb
Hydrogen	**H**	Thallium	Tl
Indium	In	Thorium	Th
Iodine	**I**	Thulium	Tm
Iridium	Ir	**Tin**	**Sn**
Iron	**Fe**	Titanium	Ti
Krypton	Kr	Tungsten	W
Lanthanum	La	Uranium	U
Lead	**Pb**	Vanadium	V
Lithium	Li	Xenon	Xe
Lutecium	Lu	Ytterbium	
Magnesium	**Mg**	(Neoytterbium)	Yb
Manganese	Mn	Yttrium	Yt
Mercury	**Hg**	**Zinc**	**Zn**
Molybdenum	Mo	Zirconium	Zr

more than a dozen. *All known substances, natural or artificial, — whether they exist in the atmosphere, as part of the earth's crust, in natural waters, as parts of plants or animals, or indeed as parts of the heavenly bodies, — are composed of one or more of the chemical elements mentioned in Table 1.* It is the task of the chemist to study the composition of substances, to build up new substances from given ones, and again to tear down complex substances, thus resolving them into simpler bodies. *Whenever a change takes place in which new substances are formed, the change is called a chemical change.* So, for example, when a piece of paper is burned, entirely new substances are formed; again, when iron rusts, a new substance, the rust, is produced. These are examples of chemical change. The chemist, however, studies not only the new products that are produced from other substances, but he also takes into consideration the means employed to cause the chemical change to take place. Then again he studies the rate with which the change proceeds, ascertaining also whether all or only part of the substance has finally been transformed. The various conditions that affect the rate with which the change takes place, and the accompaniments of the change such as alterations of temperature, volume, evolution of light, electricity, etc., are also carefully studied. *The sum total of all these observations, systematically arranged, constitutes the science of chemistry.* It touches every walk of life, for wherever there is anything material, that is to say, wherever there is anything that we can touch, hear, see, taste, or smell, we have an object with which chemistry is concerned. In fact chemical changes are continually going on about us. The food we eat is chemically transformed into the various tissues of our bodies, and these in turn are acted upon by oxygen derived from the air we breathe, and thus oxidized into simpler substances which are in turn eliminated

by means of the breath, the bowels, the kidneys, and the skin. Thus *our very lives depend upon a series of chemical changes.* Similarly the lives of all animals and plants depend upon chemical transformations of the material they take into their bodies, the products of such chemical changes being finally eliminated as the organism lives. But chemical changes are continually in progress in inanimate things, also. Thus the soil and rocks are being acted upon continually by water and the gases of the atmosphere, new products resulting, which are more or less soluble in the water. Materials from these solutions are taken up by the rootlets of trees and other plants. Furthermore as natural waters are drunk by animals, some of these soluble substances are absorbed by the membranes of the digestive tracts of the organisms and used in building up new tissues.

Many industrial processes and common operations in the household and on the farm depend upon chemical changes, and it is the purpose of this book to study some of the more important of these. It should be stated here, however, that *whenever a change takes place which is not accompanied by a formation of one or more new substances, the change is called a* **physical change.** Thus when a piece of gold is heated, it still remains gold and there is no change except a change of temperature and of volume. Again, when a glass rod is electrified by rubbing it with a silk rag, neither the substance of the rag nor that of the glass is altered. These, then, are typical examples of physical change. *Physical changes are exceedingly numerous, and they very frequently occur without any accompanying chemical change.* On the other hand, *chemical changes practically never occur without accompanying physical changes,* for the formation of new substances is generally accompanied by increase or decrease of volume, rise or fall of temperature, and frequently by evolution of light, electricity,

or still other forms of energy. So when a piece of magnesium is burned in the air, a white powder, magnesium oxide, is formed, and this is more bulky than the original magnesium; furthermore, during the burning of the metal both light and heat are evolved.

A glance at Table 1 shows that *the elements may be roughly divided into two great classes, the metals and the non-metals.* It is, however, not possible to make a perfectly sharp division into these two classes, for there are some elements that exhibit properties which partake of the nature of both the metal and the non-metal. Thus it would be easy to class such elements as copper, silver, iron, gold, mercury, zinc, and sodium as typical metals; and again, it would be simple to place such elements as sulphur, oxygen, carbon, chlorine, and phosphorus as non-metals. But elements like arsenic and antimony, for example, while possessing metallic luster and also some of the chemical properties of typical metals, are nevertheless brittle and in their chemical behavior also otherwise closely related to the non-metals. *Such elements as partake of the nature of both the metals and non-metals are sometimes called metalloids.* They are in reality a transition between the metals and the non-metals.

When a substance consists of two or more elements, it is called a **compound.** Thus oxide of magnesium is a compound, for it is formed by the union of oxygen and magnesium. Lime is a compound, for it contains the metal calcium and oxygen. Sugar is a compound, for it consists of oxygen, hydrogen, and carbon. Water is a compound, for it is composed of oyxgen and hydrogen. As *water is the most common compound on the earth,* and as the elements that compose it are also of vast importance in all plant and animal life, the chemistry of water will be our first object of study.

QUESTIONS

1. What is science?
2. What is a fact?
3. What is a theory?
4. Mention the things which the ancient Greeks called elements.
5. What is a chemical element as regarded at present?
6. About how many chemical elements are there, and how may they be classified?
7. What is a compound? Give three examples.
8. What is a physical change? Give three illustrations of physical changes.
9. What is a chemical change? Give four examples of chemical change.
10. Do chemical changes occur with or without accompanying physical changes? Explain your answer by means of a specific illustration.

CHAPTER II

THE COMPOSITION AND USES OF WATER

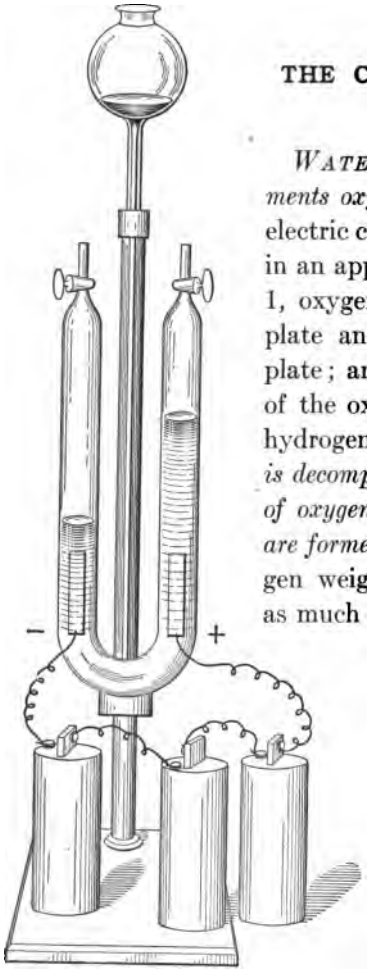


FIG. 1. — The electrolysis of water.

WATER is a compound of the elements oxygen and hydrogen. When an electric current is passed through water in an apparatus like that shown in Fig. 1, oxygen is evolved on the positive plate and hydrogen on the negative plate; and it is found that the volume of the oxygen is one half that of the hydrogen. That is to say, *when water is decomposed by electrolysis, one volume of oxygen and two volumes of hydrogen are formed.* Now one volume of oxygen weighs very nearly sixteen times as much as an equal volume of hydrogen; consequently the volume of oxygen produced in the electrolysis of water weighs eight times as much as the two volumes of hydrogen that are liberated.

By taking two volumes of hydrogen and one volume of oxygen, mixing them, and igniting the mixture with a flame or an elec-

tric spark, as, for example, in the apparatus in Fig. 2, water is formed and no hydrogen or oxygen is left uncombined. Thus the qualitative and quantitative composition of water is clearly established.

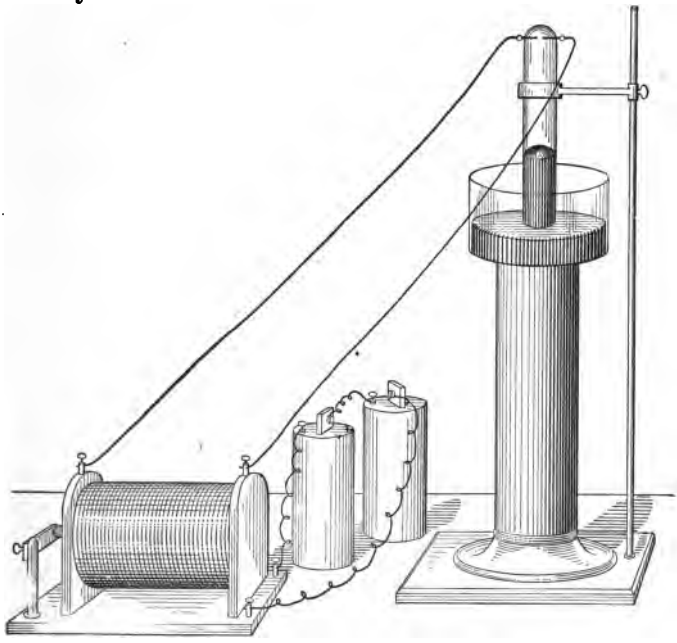


FIG. 2. — Explosion of a mixture of hydrogen and oxygen.

Though water is extremely abundant, it never is found in the pure state in nature. The waters of the oceans and many inland seas are salty. Lake water and the waters of brooks and rivers, as well as those of springs and wells, always contain more or less solid saline material which has been dissolved by the water while in contact with the soil and rocks. *On evaporating such saline waters, all of the solids remain behind as a residue.* Thus when sea water or well water is placed in a dish and allowed to evaporate either slowly at

ordinary temperature or upon boiling, there remains a residue of the material that was dissolved in the water. By evaporating several gallons of water one may obtain a sufficient amount of such dissolved matter to make a chemical analysis of it and thus determine its exact nature. Thus the dissolved matters in many terrestrial waters have been analyzed. *Both the nature and the amounts of the saline matters contained in a water are determined by the character of the rocks over which the water has coursed. The purest natural water is rain water.* It contains no mineral matter, except such small quantities of dust as have been carried into the air by the wind and then washed down by the descending drops. After it has rained awhile and the air has been well washed, the rain that falls consists of much purer water, but it always contains air dissolved in it and not infrequently nitrogenous compounds like nitrites and nitrates of ammonium, which are formed as the lightning flashes through the air. *By boiling water and condensing the vapors a fairly pure water, commonly called* **distilled water**, *is obtained.* Figure 3 shows a simple apparatus for producing distilled water. Every distilling apparatus consists of a retort, a condenser, and a receiver. In Fig. 3 these are indicated by *A*, *B*, and *C*, respectively.

The heat of the sun causes water to evaporate continually from the surface of the oceans, lakes, rivers, and moist soil. This moisture is then condensed on reaching cooler layers of air, forming clouds, fogs, dew, and rain. The rain water permeates the soil, dissolves portions of it, and gradually makes its way into brooks, rivers, lakes, and the sea. Then evaporation again goes on from the surfaces of these bodies of water, condensation again takes place, rain falls, and so the ceaseless round continues as long as the sun furnishes the needed energy. Thus it is clear that all water power is really derived from the sun. But not all of the water is thus

evaporated or carried into the sea. Some of it, though, to be sure, but a relatively small portion, is taken from the soil by the rootlets of plants, and again a certain amount is drunk by animals. *Without water, plants and animals cannot live.* Indeed *the bodies of all plants and animals contain from fifty to ninety per cent of water by weight.* This water, to be sure, is more or less firmly combined with the other materials that enter into the composition of the plant and animal bodies. From

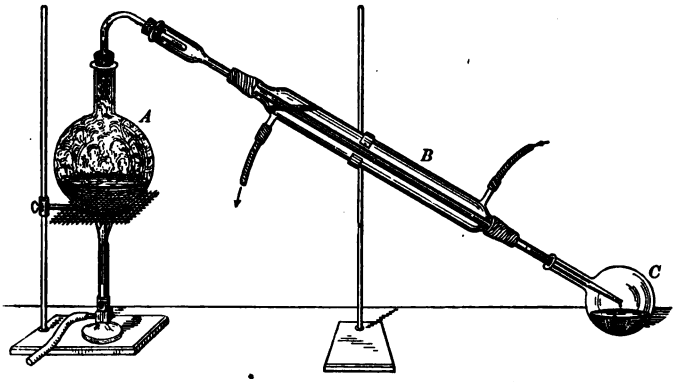


FIG. 3. — The distillation of water.

the latter it may be obtained by desiccation. That the tissues of plants and animals, soils, and not infrequently minerals, rocks, and various artificial salts like blue vitriol, epsom salts, etc., contain water may readily be shown by gently heating a small portion of the substance, about the size of a small nut, in a test tube as shown in Fig. 4. The water set free condenses in drops on the upper cooler portions of the test tube. In such experiments one soon discovers that some compounds retain the water much more tenaciously than others, for while very slight warming suffices to liberate the water in some instances, other substances need to be heated even to redness before the water is set free. So, for

example, when grass, a piece of carrot, or lean meat is gently heated in a test tube, water is readily formed in drops on the upper cooler parts of the tube. To drive the water out of an old piece of mortar or concrete would, however, require a much higher temperature. Indeed, it would be necessary to heat the substance to dull redness.

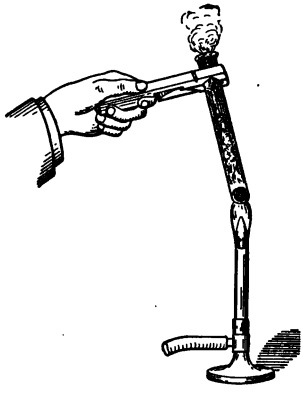


FIG. 4. — Testing a substance for water.

The amount of water contained in plants and animals is usually high. But different organisms frequently contain quite different quantities of water, the amount of which is commonly found by weighing the plant or animal substance and then drying it to constant weight at the boiling point of water, namely, 100° C. The difference between the original weight and that of the dried residue represents the moisture content. In this way it has been found that the leaves of herbaceous plants contain from 60 to 80 per cent water; potatoes, juicy fruits, and succulent plants contain from 85 to 95 per cent water; and algæ and other aquatic plants may even contain as much as 98 per cent. On the other hand, wood commonly contains only from 44 to 55 per cent water, and many dry seeds contain much less. Dry grass seed, for example, contains only about 15 per cent water. The dried residue of the plant substance may be burned, whereupon there remains *the ash, which represents the mineral matter in the plant. The combustible portion, which has been destroyed, is the so-called organic matter of the plant.* It is oxidized and largely volatilized in the process of burning

(see combustion), there being formed gaseous substances like carbon dioxide, water vapor, nitrogen, and ammonia. *The amount of ash found in different plant substances also varies considerably.* Computed on the dried material, for example, oak wood contains about 0.48 per cent ash, rye straw 4.46 per cent, rye grain 2.09 per cent, tobacco leaves 17.16 per cent, potato leaves 8.58 per cent, potatoes 3.79 per cent. *In general the leaves of plants contain more ash than the other parts.* The ash of different plants has different composition, which fact will be discussed more fully later.

The water content of animals also varies considerably. The following table indicates the amount of water contained in a few of the common animals, the entire body being considered:

TABLE 2

Animal	Water Content
Fat ox	45.5 per cent
Fat calf	63 per cent
Fat pig	41 per cent
Lean pig	55 per cent
Fat sheep	43 per cent
Lean sheep	51 per cent
Chicken flesh	74 per cent
Goose flesh	42 per cent
Turkey flesh	55 per cent
Brook trout	78 per cent
White fish	70 per cent
Lobster (exclusive of shell)	79 per cent
Oyster (exclusive of shell)	81 per cent

Table 3 presents the water content of various parts of an ox.

TABLE 3

Part	Water Content
Brain	80 per cent
Heart	63 per cent
Muscle of the loin	62 per cent
Kidney	76 per cent
Liver	71 per cent
Sweet breads	71 per cent
Lungs	80 per cent

From these data it is apparent that considerably over half of the weight of animal tissues consists of water. It should be borne in mind, however, that this water is more or less tightly combined with the other constituents of the tissues. By expelling the water in the process of drying, these combinations are ruptured, the water being thus liberated.

Now as human beings subsist on food that comes from animals and plants, it is clear that the human body derives from such food a considerable amount of the water it needs. This, however, is not sufficient, for drinking water is required in addition. *Chemically pure water is not good drinking water.* Freshly distilled water, obtained by careful distillation of spring or artesian well water, is insipid and flat to the taste, chiefly because by the process of boiling during distillation the air which was dissolved in the water has been expelled. For the same reason *water that has been boiled and then cooled without sufficient opportunity to become thoroughly aërated again has a flat taste.* Many natural spring waters and well waters are excellent drinking waters, though they always contain solid matter in solution that comes from the rocks and soils over which the water has flowed. The amount of this mineral matter varies very greatly with the nature of the solid material with which it has been in contact. So, for instance, *water that has coursed over granite rocks only has but very little mineral matter in it, for granite is but slightly soluble in water.* On the other hand, *water in limestone regions is highly charged with mineral matter, for limestone is relatively readily soluble in water especially when the latter contains carbon dioxide,* and this is always the case with water that has been in contact with the air, for the latter contains carbon dioxide (see air).

Table 4 gives the amount of mineral matter contained in a few typical natural waters in grams per 1000 liters. These

figures are obtained by evaporating a given weight of water to dryness, weighing the residue, and then computing the amount of the latter per 1000 liters of the water taken.

TABLE 4

Water from	Mineral Content in Grams per 1000 Liters
Atlantic Ocean	35,664
Indian Ocean	35,525
White Sea	33,118
Dead Sea	253,016
Great Salt Lake	302,122
Lake Michigan	145
Vichy Springs	6,798
Karlsbad Springs	6,091
Artesian well at St. Petersburg	3,891
Artesian well at London	834
Artesian well at Madison, Wis.	371
Spring near Wausau, Wis.	64
Rhine River	231
Nile River	142

It is to be borne in mind that *the mineral matter dissolved in river waters often varies materially according to the time and place of taking the sample.* The mineral content of waters of lakes and oceans also varies somewhat with the depth and locality. The figures in Table 4 are consequently to be regarded as subject to certain fluctuations. They serve, however, to give a general idea of the amount of mineral matter found in natural waters.

The mineral content of oceanic water and of salt lakes and seas consists chiefly of common salt, and such water is, of course, not fit to drink. *Spring and well waters that contain considerable amounts of calcium salts (also often called lime salts) in solution are called* **hard waters.** They may be quite good for drinking purposes, but they are not good for washing or steam boiler purposes, for in the boilers they deposit a hard coating or scale, like that commonly found in

an ordinary teakettle, for example, and with soap they form no suds, but simply a curd, which is in reality an insoluble calcium soap resulting from the decomposition of the soluble soap employed, by the calcium salts in the water. Before such water can be used for washing or boiler purposes, the lime salts it contains must be removed. This can be done in one of two ways: (1) by distilling the water, a process which on account of the fuel required is expensive and so practically out of the question for ordinary purposes, and (2) by throwing the lime salts out of solution by adding some other ingredients which will form insoluble compounds with the lime in the water. Soluble salts which thus act on the lime salts and form precipitates with them are often spoken of as **water cleansers** or water purifiers. Among those commonly in use are sodium carbonate (also called washing soda), borax, and sodium phosphate. *Of all of the natural waters rain water is by far the best and cheapest for washing and boiler purposes.* It should be collected and stored in clean tanks or cisterns to which air has ready access. By the use of Portland cement (which see) it is now an easy matter to construct suitable reservoirs for the storage of rain water at a moderate cost. Such *cistern water when clean and well aërated is good drinking water,* though its mineral content is quite low, being simply due to the small amount of material dissolved from the cement-lined walls of the cistern.

Sometimes spring water or well water is found to be unfit to drink because of the poisonous or otherwise injurious mineral ingredients it contains. But such cases are after all quite rare, being in general confined to excessively alkaline or strongly saline waters. For example, near copper deposits the waters may contain a considerable amount of compounds of that metal and so be injurious to health. In any case *when water has a taste or smell that causes suspicion, it ought*

to be properly tested to ascertain whether it is potable. Water is more frequently rendered unfit to drink because of contamination with plant or animal refuse, sewage, disease-producing

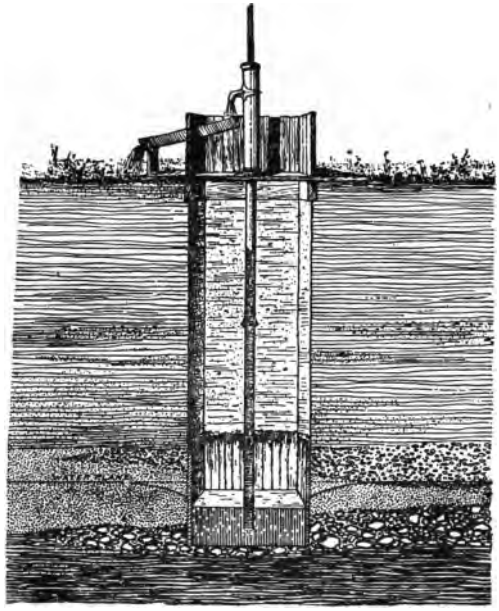


FIG. 5 (*a*). — The kind of a well into which surface water will readily seep.

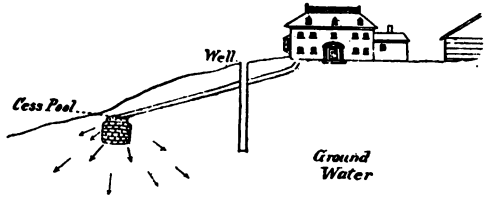


FIG. 5 (*b*). — An example of how pollution may occur.

bacteria, or other pathogenic organisms. *Wherever there is decaying animal or vegetable matter there microörganisms abound, and water contaminated with such material is dangerous*

to health. A drinking water ought consequently to be examined as to its content of microörganisms as well as subjected to chemical tests before being pronounced safe for drinking purposes. Surface wells near human habitations soon become contaminated from seepage from outhouses, cesspools, barnyards, and the like. It is to be borne in mind that after all *a well is merely a hole in the ground into which water from the surroundings tends to drain.* Mounding up the earth around the pump or opening of the well, of course, does not prevent the water of the soil in the neighborhood from getting into the well. Such a mound at the opening of the well can at best only prevent the water that is on the surface of the ground from running into the mouth of the well. *Typhoid fever is one of the commonest and most dangerous diseases that may result from drinking polluted water.* This disease is due to the *bacillus typhosus*, which is not infrequently present in contaminated water. Dysentery and other intestinal and gastric disturbances may result from the presence of other organisms. *If there is no alternative but to use water that is known to be contaminated for drinking purposes, it should first be boiled.* This destroys the organisms it contains, but the water after boiling tastes flat as already mentioned. Contaminated water is also frequently filtered through unglazed porcelain. Such filters are commonly known as the **Pasteur**

FIG. 6. — Typhoid bacillus greatly magnified.

water filters. They remove nearly all of the matter suspended in the water as well as a very large share of the organisms that are present. But such filters need to be renewed frequently, for the refuse they accumulate in their pores soon becomes only an added source of danger toward further contamination. On a large scale water is frequently purified by filtration through properly constructed filters of gravel or crushed rock and sand. These too are fairly efficient in removing dangerous material suspended in the water, but they also require to be renewed from time to time. Water derived from wells that are a hundred feet or more deep has obviously been well filtered in coursing through the various strata to the depth from which it issues, and consequently such water is practically always free from organic contaminations. *Deep well water is therefore usually to be preferred for drinking purposes.*

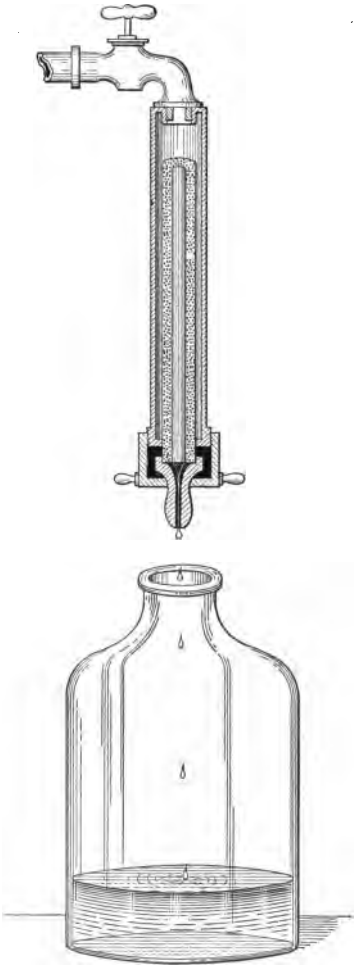


FIG. 7. — Pasteur water filter.

Besides resorting to boiling or filtration for the purpose of

removing organic contamination from a potable water supply, chemical means of purification may be adopted. This always involves introducing some chemical substance into the water, which destroys the organisms it contains. For such purposes a solution of sodium or calcium hypochlorite (bleaching powder) is not infrequently employed, especially in the case of a city water supply where the water in the reservoir and mains shows a high bacterial content. It is best to avoid the introduction of such chemicals into drinking water, for these substances which destroy bacteria are also deleterious to health. It is only when the disease germs are actually present in the water that such chemical treatment is really justifiable, for the germs are far more dangerous than the relatively small amounts of the antiseptics required to kill them. *An antiseptic is a chemical that destroys germs, or prevents or retards their growth, without exerting a materially harmful effect upon the living body. Antiseptics therefore frequently are germicides.* These will be referred to again.

QUESTIONS

1. How may it be shown that water is a compound of oxygen and hydrogen?

2. In what proportions by weight do these elements occur in water? In what proportion by volume?

3. Why are natural waters not chemically pure? What impurities are there in rain water? In well water?

4. Draw a diagram of an apparatus for making distilled water, naming the principal parts of the apparatus.

5. Make a list of the important ways in which water is found in nature.

6. How may one determine whether a substance contains water? Give a specific example.

7. About how much water is there in an animal like a cat? In a dahlia root?

8. What are some of the important characteristics of a good drinking water?

9. What is meant by the term hard water?

10. How would you treat a water that is known to be polluted and yet must be used for drinking purposes? Why?

CHAPTER III

HYDROGEN, OXYGEN, HYDROGEN PEROXIDE, AND OZONE

WHILE both hydrogen and oxygen are obtained when the electric current is passed through water, these gases may also be prepared by other methods. On account of the frequent occurrence of these elements in many common substances, it is necessary to study their properties somewhat more closely.

Hydrogen is commonly prepared by the action of an acid, like hydrochloric or sulphuric acid, on certain metals, like zinc or iron, for example. It is a colorless, odorless, tasteless gas, which is very light, in fact the lightest of all known gases, being 14.388 times lighter than air. One liter of pure dry hydrogen at 0° C. and under atmospheric pressure at sea level weighs 0.08987 gram. Hydrogen is an inflammable gas. When pure it may be burned from a jet. The flame is almost colorless, but very hot. Mixtures of air and hydrogen, or oxygen and hydrogen, are explosive, and great care must be taken not to heat such mixtures or to ignite them with a flame or a spark. By cooling hydrogen to − 253° C. it may be liquified at atmospheric pressure. Liquid hydrogen is clear and colorless, that is, like water in appearance, but it is only 0.07 as heavy as water. At − 259° C. hydrogen solidifies, forming white crystals. Hydrogen gas is not poisonous, but animals and plants would die from suffocation when kept in hydrogen, because they require oxygen to breathe. Be-

cause of its lightness hydrogen is used for filling balloons. Though the gas burns with quite a hot flame, it is not used as a fuel in the pure condition, for this would be too expensive. *Hydrogen is, however, one of the constituents of coal gas, natural gas, and water gas, all of which are used as fuels. When hydrogen burns in the air or in oxygen, the product formed is water.* When steam is passed over red-hot iron, hydrogen is formed and black oxide of iron, hammer black, is simultaneously obtained, for at the high temperature the iron

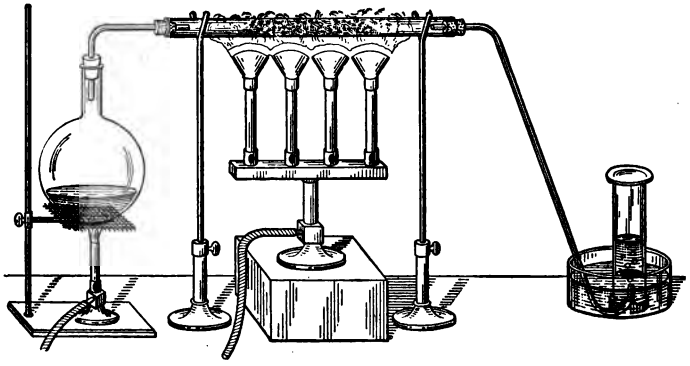


FIG. 8. — Making hydrogen by passing steam over red-hot iron.

unites with the oxygen in the water and thus the hydrogen is set free. Magnesium will similarly liberate hydrogen, and in this case it suffices to immerse that metal in boiling water. Sodium will liberate hydrogen from water rapidly even at room temperature. Potassium acts even more vigorously than sodium. In the case of sodium and potassium, the oxides formed pass into solution in the excess of water and form dilute lye. These solutions turn red litmus paper blue, are alkaline to the taste, and feel slippery to the touch. Stronger solutions of lye are caustic and disintegrate the flesh.

While hydrogen practically does not occur in nature in the free, i.e. *uncombined, condition, it is exceedingly abundant as a constituent of many compounds.* As already mentioned, one ninth of the weight of water is hydrogen. *All plant and animal tissues contain hydrogen. Crude petroleum and all of its various products contain hydrogen.* This element is also found in many rocks.

Oxygen *occurs in the free state in the air.* In fact it forms about 21 per cent of the total volume of the air. Combined with other elements, oxygen is also found in large quantities and very widely distributed over the earth. In fact it is the most abundant of all of the elements, which is readily apparent, for *88.88 per cent of the weight of all water is oxygen, and the latter also forms 44 to 48 per cent of the weight of the solid portions of the earth. In all plants and animals oxygen is combined with hydrogen, carbon, nitrogen, phosphorus, and minor amounts of other elements.*

Pure oxygen is most readily prepared by heating certain compounds which contain it, and which give up their oxygen content fairly readily at higher temperatures. The substance which is most frequently chosen for this purpose is potassium chlorate. It consists of white crystals whose constituents are potassium, chlorine, and oxygen; the latter element forms 39.1 per cent of the weight of this salt. Thus from 100 lb. of potassium chlorate 39.1 lb. of oxygen may be obtained. The oxides of mercury and silver will also decompose on heating, yielding free oxygen, and metallic mercury and silver respectively. Again, saltpeter when heated will give up a portion of its oxygen. *Plants are continually giving off oxygen as they breathe.*

Pure oxygen is colorless, odorless, and tasteless. It is 1.1 times as heavy as air and sixteen times as heavy as hydrogen. A liter of oxygen at 0° and under a barometric

pressure of 760 mm. (*i.e.* under the so-called standard conditions of temperature and pressure under which gases are measured) weighs 1.429 grams. Liquid oxygen is pale blue in color, boils at − 182°.5 C., and is 1.1315 times as heavy as water. It is attracted by a magnet. By chilling liquid oxygen sufficiently, snow-white crystals may be obtained, which melt at − 227° C. In water oxygen is but slightly soluble, 1 volume of water absorbing but 0.034 volume of oxygen at 15° C. This slight solubility of oxygen in water enables us to collect the gas over water. *The fact that oxygen is soluble in water, even if but slightly, is nevertheless very important, for fishes and other organisms living in water depend upon this dissolved oxygen for their supply.*

In its ability to form compounds with other elementary substances oxygen stands at the head of the entire list of the chemical elements. In fact, it will unite with all of the elements except fluorine and the members of the so-called argon group, namely argon, helium, neon, krypton, and xenon. *All ordinary combustion in the air is really union of the burning substance with oxygen.* That is to say, ordinary combustion is oxidation. So when coal, petroleum, paper, or wood are burned in the air they are oxidized. A part of the products formed is gaseous and so escapes in the air, and another portion remains behind in the solid state as ash. Now, when substances are burned in pure oxygen, the process goes on far more rapidly and more vigorously. In fact, some substances will burn brilliantly in pure oxygen, whereas in the air they either would not burn at all or the action would proceed very slowly, for it must be borne in mind that the oxygen in the air is diluted, so to say, with four times its volume of nitrogen, which is a rather inert gas and does not take part in ordinary combustion. Heated charcoal glows brilliantly in oxygen. Sulphur and phosphorus burn brilliantly

in oxygen gas, and even the steel of a watch spring burns with brilliant scintillations in that gas. One can breathe pure oxygen for a time without bad effects; in fact, the gas is frequently administered to patients who experience difficulty in breathing and so get an insufficient oxygen supply when breathing air. It is inadvisable, however, for a normal person to breathe oxygen, for thus the system gets too great a supply of that substance and the oxidation processes in the system go on too rapidly.

FIG. 9. — Burning phosphorus in oxygen.

The oxides of carbon, phosphorus, sulphur, and certain other elements readily dissolve in water and form solutions that are sour to the taste and turn blue litmus red. Such solutions are called acid solutions. It was at one time thought that all acids contain oxygen; this is, however, not the case, though it is true that by far the larger number of all known acids do contain oxygen. The word oxygen means acid generator. It has been retained, even though all acids do not contain oxygen. The fact is that all acids do contain hydrogen, which will be considered later.

Some oxides, for example, *the oxides of sodium, potassium, and calcium* (obtained by burning each of these metals in

oxygen) *are white powders which dissolve in water and yield solutions that are alkaline to the taste, slippery to the touch, and turn red litmus blue.* Such solutions are called alkaline solutions. When acid solutions and alkaline solutions are mixed, they mutually neutralize each other. The product thus formed is a salt, which, remaining in solution, has neither an acid nor an alkaline taste, and does not affect either red or blue litmus (see acids, alkalies, salts). It thus appears that by union with some elements acidic oxides or acid-forming oxides result, and by union with other elements alkaline oxides or alkali-forming oxides are obtained. It is true also that in some cases oxides are formed which are neither acid nor alkali forming, but rather indifferent or neutral bodies, like the oxides of iron, copper, and mercury, for example.

Ozone may be prepared from oxygen. Three volumes of oxygen yield two volumes of ozone. On heating the latter ordinary oxygen is formed again, three volumes being obtained from two volumes of ozone. Ozone is then 1.5 times as heavy as oxygen. Ozone is produced when pure oxygen or the air is subjected to the action of electric sparks. So, for instance, in the neighborhood of a frictional electrical machine in action there is always a certain amount of ozone produced, readily distinguished by its peculiar garlic-like odor. This odor of ozone is also observed when lightning strikes terrestrial objects or when phosphorus slowly oxidizes at room temperatures in moist air. *Ozone is the most powerful oxidizing agent known and this is its most important property.* Ozone will not only kill germs and other microscopic organisms rapidly, but it will also oxidize dyestuffs, destroying their color, and convert silver, lead, arsenic, sulphur, and other elements into oxides. On account of its great activity, ozone exists in the air for but a short time after it has been formed. Ozone acts on water, forming **hydrogen peroxide,**

also called **hydrogen dioxide.** This is a compound of oxygen and hydrogen which contains just twice as much oxygen as does water. Thus in water 1 gram hydrogen is united to 8 grams of oxygen, while in hydrogen peroxide 1 gram of hydrogen is united to 16 grams of oxygen. In the pure state hydrogen peroxide is a thick sirupy liquid whose specific gravity is 1.458. It is colorless, but like water it looks blue in thick layers. This liquid is quite unstable and decomposes readily into water and oxygen. In dilute solutions it is more stable, especially when kept cool and in the dark. It is now a common article of commerce, the 3 per cent solutions frequently being sold under the trade name of "dioxygen." This solution is valuable as a mild bleaching agent and as an antiseptic. Its power as a germicide depends upon the fact that it readily decomposes into oxygen and water. The oxygen set free destroys the organisms. Hydrogen peroxide is commonly prepared by the action of cold dilute sulphuric acid upon barium peroxide, thus:

Barium peroxide plus sulphuric acid yields hydrogen peroxide plus barium sulphate.

The latter substance may be filtered off. The clear filtrate contains the hydrogen peroxide in solution. On account of the fact that hydrogen peroxide also results when water is acted upon by ozone, the latter cannot exist long in the air, which always contains moisture, and this would interact with ozone, forming hydrogen peroxide and oxygen.

QUESTIONS

1. What are the important properties of hydrogen?
2. How may hydrogen be prepared?
3. In what compounds is hydrogen found in nature?
4. What is formed when hydrogen burns? How may this be shown?

5. How does oxygen occur in nature?

6. How may pure oxygen be made?

7. How many pounds of oxygen could be prepared from 83 pounds of potassium chlorate?

8. Mention the important properties of oxygen.

9. What is an oxide? How may oxides of phosphorus, carbon, sulphur, and iron be formed?

10. How may ozone be formed? Mention its important properties. What is ozone?

11. What is hydrogen peroxide? How may it be prepared?

12. Mention the uses of hydrogen peroxide.

CHAPTER IV

THE AIR, NITROGEN, NITRIC ACID, AND AMMONIA

THE main constituents of the air are nitrogen and oxygen. These are present in about the proportions of 78 volumes of the former to 21 volumes of the latter. Besides these two gases there are present, however, water vapor, carbon dioxide, and the elements of the argon group, namely helium, neon, krypton, and xenon. Table 5 gives the composition of a sample of air as it is found in the country or over the sea.

TABLE 5

COMPOSITION OF A SAMPLE OF NORMAL COUNTRY AIR

100 volumes contain as follows:

Nitrogen	77.42 volumes
Oxygen	20.77 volumes
Argon, helium, neon, krypton, and xenon	0.93 volume
Water vapor	0.85 volume
Carbon dioxide	0.03 volume
	100.00 volumes

The amount of moisture which air contains varies considerably from time to time, depending especially upon temperature and proximity to bodies of water. Minor quantities of sulphur dioxide, hydrogen, ammonia, nitric acid, particles of dust, and various microbes are also present in air. The amounts of all of these are also quite variable. While the members of the argon group are present in air to the extent of over 0.9 per cent by volume, these gases are nevertheless of no practical importance and hence will receive no further con-

sideration here. The carbon dioxide in the air in the country or over the sea amounts to about 0.03 per cent by volume, but in cities where much fuel is consumed the air often contains double that amount and even more, while in densely crowded audience rooms the carbon dioxide content may run as high as six to eight times that found in city air. *While the carbon dioxide in the air is present to the extent of only about 0.03 per cent, yet it must be borne in mind that plants get all of their carbon from the carbon dioxide in the atmosphere.* They breathe in carbon dioxide, which in the presence of sunlight in the green leaf of the plant interacts with the moisture that is present, forming starch and setting oxygen free, which the plant therefore exhales. This change may be written as follows:

Carbon dioxide + **water** (in the sunlight in the green leaf of the plant) **yields starch** + **oxygen.**

The constituents of the air, Table 5, are not chemically united; they form merely a mixture. In spite of this the relative amounts of oxygen, nitrogen, and argon are nearly constant in the atmosphere everywhere.

The highest amounts of moisture which the air is able to hold at different temperatures is shown in Table 6.

TABLE 6

Temperature of Saturation	Amount of Water Vapor in 1 Cu. Meter of Saturated Air
−5° C.	4.0 grams
0° C.	5.4 grams
+5° C.	7.3 grams
10° C.	9.7 grams
15° C.	13.0 grams
20° C.	17.1 grams
22° C.	22.5 grams
30° C.	30.0 grams

If the air has but 0.4 of the water vapor in it that it can hold, we feel that the air is dry. If, however, 0.8 of its maximum capacity of moisture is present in the air, we feel that it is moist or humid. As the moist air reaches the upper and cooler regions of the atmosphere, partial condensation of the water to minute drops takes place, forming mists, which on account of their altitude are commonly called clouds. When the mists form near the surface of the earth, they are called fogs. These are formed when moisture-laden air is cooled to such an extent that a portion of its moisture is condensed. This is frequently caused by a cold current of air striking warmer moisture-laden strata. When the condensation in the upper regions of the air is such that large drops are formed, these fall to the ground as rain, and carry with them in solution some of all of the gases in the atmosphere, together with fine particles of dust, microörganisms, etc., that were suspended in the air. This material on reaching the ground soaks into the soil in part, and in part it flows off the surface into brooks, rivers, lakes, and the sea. It has been estimated that about 80 per cent of the water that falls as rain on the land soaks into the ground and 20 per cent runs off the surface into the waterways. Snow, sleet, and hail form at temperatures at and below the freezing point of water. Dew forms at night. It consists of drops of moisture that deposit upon objects which cool off relatively rapidly, and so chill the moisture-laden air that comes into contact with them. The importance of these various aqueous precipitations for the life of plants will receive further consideration later.

The common way of preparing nitrogen gas is to abstract the oxygen from an inclosed space of air. This may be done in various ways. So, for example, an animal like a rat or a

mouse may be placed in a tightly stoppered bottle, whereupon it will soon suffocate, because it will use up the oxygen in the air in the bottle and breathe out carbon dioxide. The latter gas may be removed by shaking the gas left in the bottle with some clear limewater. The limewater becomes milky in appearance because a precipitate of carbonate of lime, calcium carbonate, is formed, thus:

Limewater + carbon dioxide yields
calcium carbonate + water.

By this chemical change, then, the carbon dioxide is removed, being made part and parcel of a solid substance, the calcium carbonate, which is chemically the same as chalk. Now the gas which still remains in the bottle is nitrogen plus the gases of the argon group. These latter in all make up only about 0.9 per cent of the entire air. They have all been discovered in the air since 1895. They are very inert, in fact thus far it has been found to be impossible to get them to unite chemically with any other element. *By simply removing the oxygen and carbon dioxide from the air, pure nitrogen is not obtained, for the gases of the argon group are still present.* For the present purposes, however, it will not be necessary to devote space to the description of how the latter gases may be removed; suffice it here to say that perfectly pure nitrogen is best prepared from certain pure chemical compounds in which it occurs. The oxygen of the air may also be removed by burning certain substances in the air in a closed space. So one may burn a bit of sponge

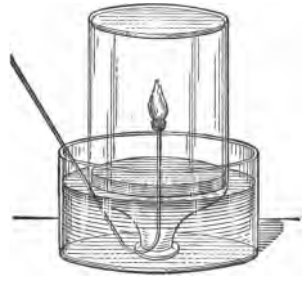


FIG. 10. — Removing oxygen from the air by burning alcohol in it.

saturated with alcohol in a bottle inverted in a dish of water as shown in Fig. 10. In this case water and carbon dioxide are formed and the latter gas is absorbed by the water in the bottle. In place of the alcohol a bit of phosphorus may be burned similarly in the bottle. In this case the phosphorus unites with the oxygen, forming phosphoric oxide, which dissolves in the water, thus leaving the nitrogen intact. In both of the last experiments the water rises on the inside of the bottle, filling the latter to the extent of one fifth of the original volume occupied by the air, thus showing that *four fifths of the air is nitrogen and one fifth oxygen.*

If a burning splinter or a lighted candle is thrust into nitrogen gas, the flame is at once extinguished, showing that *nitrogen will not support combustion.* Animals suffocate in nitrogen, and die for lack of oxygen. Nitrogen gas is not poisonous, but it cannot be used in respiration. In the air it dilutes the oxygen that is present and so retards the too rapid oxidation that would take place if pure oxygen were inhaled. *At ordinary temperatures nitrogen is quite an inert gas, but at high temperatures it does combine with other substances chemically.* So, for example, nitrogen does not unite with oxygen when these gases are mixed. In fact no union takes place, even if these gases are heated together quite highly, as, for instance, when the mixture is passed through a red-hot tube. But if a mixture of nitrogen and oxygen is subjected to the very high temperature of the electric arc, union does take place, oxide of nitrogen being thus formed. The latter when treated with air and water yields nitric acid, and indeed this process is now used successfully in forming that important substance on a commercial scale. As electricity is required in the process, this method of manufacturing nitric acid is only profitable where water power is

available for the production of cheap electric energy. In Norway considerable quantities of nitric acid and nitrates, which are important as fertilizers, are thus manufactured at present. As the lightning flashes through air during thunderstorms, nitric acid and nitrates are similarly produced, and these as they are carried down by the rain enrich the soil. It is for this reason that *a thundershower is much more helpful to vegetation than merely an ordinary rain. The latter supplies water, but the former furnishes fertilizer as well as water.*

The two most important common compounds of nitrogen are nitric acid and ammonia. How nitric acid may be made

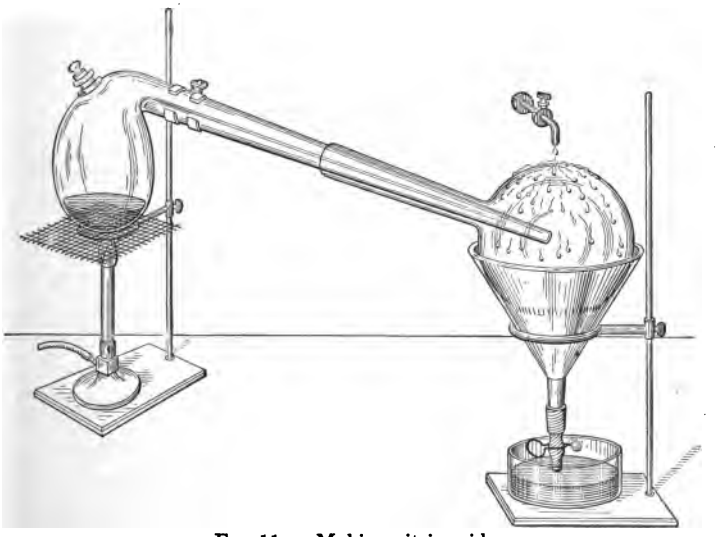


FIG. 11. — Making nitric acid.

from the air and water has just been stated. It may also be made from saltpeter, *i.e.* potassium nitrate, by treatment with strong sulphuric acid and heating. Thus the nitric acid is liberated and distills over into the receiver as shown in Fig. 11. Instead of the somewhat costly potassium ni-

trate, sodium nitrate, which is mined in Chili and consequently called Chili saltpeter, may be used. Thus nitric acid and sodium sulphate are produced, instead of nitric acid and potassium sulphate. We may write these changes thus:

Potassium nitrate + sulphuric acid =
nitric acid + potassium sulphate.

Sodium nitrate + sulphuric acid =
nitric acid + sodium sulphate.

Pure **nitric acid** is a colorless liquid having a pungent odor. Its specific gravity is 1.414 at 15° C. In the sunlight the liquid turns yellowish in color, due to partial decomposition. *Nitric acid is a very powerful acid and also a strong oxidizing agent.* It turns blue litmus red, discolors indigo solution, disintegrates cloth, corrodes the flesh, and turns the skin yellow. Metals, with the exception of gold and platinum, are acted upon by nitric acid, being converted either into nitrates or oxides. *The various salts of nitric acid are all called* **nitrates.** The nitrates, especially those of potassium, ammonium, calcium, and sodium, are very important as fertilizers. While nitric acid does not occur in nature as such, its salts, that is, the nitrates, are quite widely distributed in the soil. In the air nitrate of ammonium is also present. *Nitrates are exceedingly important as soil constituents,* for it is from them that plants get their supply of nitrogen. The nitrates in the soil are contained in the soil water which holds them in solution.

Nitric acid is also used in manufacturing sulphuric acid, in which case it serves to oxidize sulphur to its highest stage of oxidation. Furthermore, in making collodion, gun cotton, dynamite, and smokeless powder, nitric acid is used in large quantities.

Nitrous acid may be obtained from nitric acid by robbing the latter of a part of its oxygen. *The salts of nitrous acid are called* **nitrites**; they occur in the soil, as a result of the decomposition of plant and animal matter. *All plants and animals contain nitrogen in combination with carbon, hydrogen, oxygen, and minor amounts of sulphur and phosphorus.* When the plant and animal tissues decay, nitrogen is set free in part, and in part it is converted into ammonia, a compound of nitrogen and hydrogen, and the ammonia is gradually oxidized to nitrites and finally to nitrates. The latter represent the highest oxidation products. *While the nitrates in the soil are a very important source from which plants derive their supply of nitrogen, they are by no means the only source.*

Ammonia *is a compound of nitrogen and hydrogen.* It contains 14 parts of the former to 3 parts of the latter by weight. By volume the composition is represented by the following equation:

3 volumes of hydrogen + 1 volume of nitrogen = 2 volumes of ammonia.

Hydrogen and nitrogen do not combine at all readily, however, even when the gases are heated together highly, as, for instance, by means of the electric spark. On thus subjecting a mixture of hydrogen and nitrogen gases to the continuous action of the electric spark, a small fraction of a per cent of the gases present is converted to ammonia. But by heating nitrogenous plant or animal tissues, with little or no access of air, ammonia is readily obtained, especially when the organic matter used is mixed with lime or caustic soda or potash and then subjected to heat. In fact *the ammonia and all of the ammonium compounds of commerce are obtained as a by-product of the manufacture of illuminating gas from*

coal. In the latter process the coal, which represents the remains of the vegetation of the carboniferous age and other early geological periods, is heated out of contact of the air. The gases that escape are washed by gurgling them through water, and the ammonia is then found dissolved in this water. Plant and animal matter, including coal, peat, etc., all contain nitrogen and hydrogen in combination with carbon, oxygen, sulphur, phosphorus, etc., and when this material is heated out of contact with the air (*i.e.* subjected **to dry distillation**, also called **destructive distillation**), a variety of gaseous products are formed, among which is ammonia. The latter forms by union of nitrogen and hydrogen from the organic matter.

Ammonia gas is colorless and only 0.59 as heavy as air. It has a very characteristic, sharp, penetrating odor, irritates the mucous membranes, and causes a flow of tears. One cannot inhale the gas. Animals soon die in it. The gas does not support combustion and does not burn in the air; but it may be burned in oxygen. About 700 volumes of ammonia gas will dissolve in 1 volume of water at ordinary temperatures. This solution is lighter than water. At 14° C. the saturated solution has a specific gravity of 0.8844 and contains 36 per cent of ammonia by weight. The solution of ammonia gas in water is popularly called **ammonia water, spirits of hartshorn,** and **aqua ammonia.** It has the same odor as ammonia gas, for the latter is continually being given off by the solution. Indeed, all of the ammonia gas may be expelled from the solution by boiling the same. Finally water alone remains behind. Ammonia turns red litmus blue, and is consequently alkaline, which is further evidenced by the fact that it will unite directly with acids, neutralizing them and forming ammonium salts. So with nitric acid ammonia forms ammonium nitrate, with

sulphuric acid **ammonium sulphate**, with hydrochloric acid **ammonium chloride**. *All ammonium salts may be volatilized by heating them. On treating any ammonium salt with lime or lye, ammonia gas is evolved,* and this is the best way to prepare pure ammonia gas. For example:

Ammonium chloride + soda lye =
sodium chloride + ammonia + water.
(common salt)

Ammonia water is used in the household in cleansing clothes and polishing metals. It does not attack the latter as drastically as acids do, and consequently is less objectionable.

Ammonia is formed wherever nitrogenous plant and animal remains are decaying. So, for instance, it is found in stables, in dung piles, in the leachings from the latter, and in the soil. When thus formed from rotting organic matter, ammonia at once combines with any acids that may be present, especially with carbonic acid gas, which is always at hand, being a constituent of the air. *In the soil, then, ammonia is present in the form of ammonium salts, among these ammonium nitrate, ammonium nitrite, ammonium carbonate, ammonium chloride, and ammonium sulphate are the most important.* All of the ammonium salts are soluble in water, and hence are present in solution in the soil water. *From ammonium salts thus dissolved in the waters of the soil, plants derive a considerable share of their supply of nitrogen,* hence the value of ammonium salts as fertilizers. In commerce ammonium sulphate, which contains 21 per cent nitrogen, is quite generally sold as a fertilizer. It is manufactured at the gas works in large cities as a by-product in making coal gas.

QUESTIONS

1. (*a*) What are the constituents of the air?

(*b*) In what proportions do the two main constituents occur in the air?

2. Of what use is the carbon dioxide in the air?

3. When does the air feel dry? When moist?

4. What becomes of the water that falls to the ground as rain?

5. What is dew? Snow? Sleet? Hail?

6. How may nitrogen be prepared? Give its principal properties.

7. Why will an animal die in nitrogen gas?

8. How may nitric acid be formed?

9. What are the properties of nitric acid?

10. Of what use are nitrates in the soil?

11. What is ammonia? How prepare it?

12. What is ammonia water? What is it used for?

13. Why are ammonium salts useful as fertilizers? From what source are these salts obtained?

14. Mention the important forms in which nitrogen occurs in nature.

CHAPTER V

ACIDS, ALKALIES, SALTS, AND CHEMICAL FORMULAS

We have already learned that when carbon, sulphur, or phosphorus are burned and the products of combustion are dissolved in water, liquids are obtained which have a sour taste and redden blue litmus. Substances possessing such characteristics are commonly termed **acids**. *Acids occur in nature in plants, in animals, and also in the mineral world.* So, for example, the sourness of the lemon is due to the *citric acid* which it contains. The latter substance may be prepared from lemon juice. It consists of beautiful colorless crystals which are readily soluble in water. The solution has a markedly sour taste. This same citric acid is found in other citrous fruits as well, and is most abundant in them before they are quite ripe. In sour apples, mountain ash berries, and many other similar fruits *malic acid* is present. It too may be prepared from these in the pure state. It is a white solid substance which is readily soluble in water. Similarly grapes contain *tartaric acid.* The jack-in-the-pulpit contains *oxalic acid.* On fermentation of apple juice, *vinegar,* which is essentially a dilute solution of *acetic acid,* is formed. When milk sours, *lactic acid* is formed, and indeed lactic acid also occurs in the muscles of man and animals. In the human stomach hydrochloric acid is formed by the gastric glands. In red ants *formic acid* is found. The air, as we have seen, always contains a small amount of carbonic acid, and in cities where much coal is consumed, it

also contains sulphurous acid, for coal contains sulphur. All natural waters, too, contain carbonic acid in solution, and some spring waters are highly charged with this substance. Soil waters also contain acids formed during the processes of decay of animal and vegetable matter. *Boric acid* is found in volcanic regions to some extent, and *silicic acid* is one of the most abundant of all substances. It occurs as quartz in huge masses, often forming mountains, and as sand grains it covers vast areas of the earth's surface. Silicic acid is practically insoluble in water. It consequently has no taste and does not redden litmus. However, with alkalies it has the power to form salts, which indeed is after all the most important characteristic of all acids. In fact *some acids are so weak that they do not redden litmus, and have no perceptible sour taste, and yet they are unquestionably acidic substances because they do form salts with alkalies.*

Lime, as it is used for building purposes, soda lye, potash lye, and ammonia are substances which turn moist red litmus paper blue. They are typical **alkalies.** In strong solutions they have a caustic or corrosive action on the skin. With acids they react chemically, forming new substances called salts. So, for example, when a solution of soda lye is treated with a solution of hydrochloric acid till the resulting liquid does not change either red or blue litmus paper, we say that the acid and the lye have *neutralized* each other. The substance which is now in the solution is common salt. It has a salty taste, whereas the acid had a sour taste, and the lye had an alkaline taste which is quite peculiar to itself. The change which takes place when the acid and the lye act on each other may be expressed in the form of an equation thus:

Soda lye + hydrochloric acid yields
sodium chloride (*i.e.* common salt) + **water.**

Now it is somewhat cumbersome to write such equations in words, and in order to save space and time, chemists have adopted a system of abbreviations which is very easily comprehended. Each element is represented by the initial letter of its name. Thus C is the symbol for carbon, H stands for hydrogen, O for oxygen, N for nitrogen, P for phosphorus, I for iodine, etc. As several of the elements have names that begin with the same letter, their symbols are distinguished from one another by adding to the first letter (which is always capitalized) one other characteristic letter contained in the name of the element. This second letter is never capitalized. So, for example, C stands for carbon, Ca for calcium, Co for cobalt, Cl for chlorine, etc. While in some cases the symbol is derived from the common name of the element (as already illustrated), in other instances it is derived from the Latin name of the element. So sodium, *natrium*, has the symbol Na; potassium, *kalium*, K; copper, *cuprum*, Cu; silver, *argentum*, Ag; iron, *ferrum*, Fe; mercury, *hydrargyrum*, Hg; tin, *stannum*, Sn; gold, *aurum*, Au; (compare Table 1). Now these symbols do not stand for the names of the elements only, but they also represent definite quantities by weight in which the elements combine chemically. So H stands for hydrogen, but also for 1 gram of hydrogen; Cl stands for chlorine, but also for 35.5 grams of chlorine; O represents oxygen, but also 16 grams of oxygen; etc. Table 7 gives the names of the more common elements together with their symbols and the number of parts by weight for which each symbol stands. A complete list of the chemical elements and their symbols has already been given on page 3. While Table 7 mentions but thirty-eight of the elements, the data are nevertheless quite sufficient for all ordinary purposes, for the elements that have been omitted are of minor importance.

TABLE 7

Aluminum	Al	27.1	Lead	Pb	207.10
Antimony	Sb	120.2	Lithium	Li	6.94
Arsenic	As	74.96	Magnesium	Mg	24.32
Barium	Ba	137.37	Manganese	Mn	54.93
Bismuth	Bi	208.0	Mercury	Hg	200.6
Boron	B	11.0	Molybdenum	Mo	96.0
Bromine	Br	79.92	Nickel	Ni	58.68
Cadmium	Cd	112.40	Nitrogen	N	14.01
Calcium	Ca	40.07	Oxygen	O	16,0
Carbon	C	12.0	Phosphorus	P	31.04
Chlorine	Cl	35.46	Platinum	Pt	195.2
Chromium	Cr	52.0	Potassium	K	39.10
Cobalt	Co	58.97	Silicon	Si	28.3
Copper	Cu	63.57	Silver	Ag	107.88
Fluorine	F	19.0	Sodium	Na	23.0
Gold	Au	197.2	Strontium	Sr	87.62
Hydrogen	H	1.008	Sulphur	S	32.07
Iodine	I	126.92	Tin	Sn	119.0
Iron	Fe	55.84	Zinc	Zn	65.35

The figures in Table 7 have all been determined by ascertaining the weights of the elements that unite to form compounds. Experience has shown that *chemical compounds when pure always contain the same elements in the same proportion by weight.* This is called the **law of definite proportions.** So, for instance, hydrochloric acid always consists of hydrogen and chlorine, and nothing else. Moreover, in hydrochloric acid there are 35.5 grams of chlorine combined with every gram of hydrogen; that is to say, for every 1 part by weight of hydrogen, hydrochloric acid contains 35.5 parts by weight of chlorine, and this is represented by the symbol HCl, commonly termed the formula for hydrochloric acid. Again, in common salt, which is sodium chloride, we always have 35.5 parts by weight of chlorine combined with 23 parts by weight of sodium. *The symbol or formula* for common salt, then, is NaCl, and it *expresses both the qualitative and quantitative composition* of sodium chloride.

We may indeed regard common salt, *i.e.* NaCl, as derived from hydrochloric acid, *i.e.* HCl, the 1 part by weight of hydrogen of the latter being replaced by 23 parts of sodium by weight. Thus 23 parts of sodium by weight may replace 1 part of hydrogen by weight, and consequently 23 is said to be the *hydrogen equivalent* of sodium; for *it has been found to be true in general that whenever sodium can replace hydrogen in a compound it takes 23 grams of sodium to play the rôle of 1 gram of hydrogen.* The hydrogen equivalent of other elements may be found similarly. Such replacements are very easily represented by means of the chemical symbols mentioned. Thus:

$$\underset{\text{(36.5 grams hydrochloric acid)}}{\mathrm{HCl}} + \underset{\text{(23 grams sodium)}}{\mathrm{Na}} = \underset{\text{(58.5 grams common salt)}}{\mathrm{NaCl}} + \underset{\text{(1 gram hydrogen)}}{\mathrm{H}}$$

In general, then, *the composition of any compound is expressed by writing the symbols of its elements one after the other.* However, whenever the compound is a gas or vapor, the symbols that represent it indicate not only the qualitative composition and quantitative composition by weight, but also the weight of 22.4 liters of the gas or vapor at 0° C. and 760 mm. barometric pressure. So, for example, the symbol HCl stands for hydrochloric acid and indicates that in this compound 1 gram of hydrogen is combined with 35.5 grams chlorine; it shows also, however, that 22.4 liters of the gas weigh 36.5 grams, whence 1 liter weighs $\frac{36.5}{22.4}$, or 1.629 grams. The volume 22.4 liters is chosen because it is the volume of 2 grams of hydrogen at 0° C. and 760 mm. pressure, with which it is customary to compare other gases. It is not necessary to enter further into the reasons for this practice here. For our purpose it suffices to know that *in the case*

of any gas the weight of a liter of it may be found by dividing the formula weight by 22.4. The following examples will serve as illustrations. The composition of carbon dioxide gas is represented by CO_2, which indicates that 12 grams of carbon are combined with 2×16 or 32 grams of oxygen; furthermore, the formula indicates that $12 + 32$ or 44 grams (the formula weight) is the weight of 22.4 liters of the gas, whence 1 liter weighs $\frac{44}{22.4}$, or 1.964 grams. Again, the formula for water is H_2O, which shows that it consists of hydrogen and oxygen in the ratio of 2 grams of the former to 16 grams of the latter; it also indicates, however, that $2 + 16$ or 18 grams of water vapor occupy 22.4 liters of space under standard conditions, whence 1 liter of the vapor weighs $\frac{18}{22.4}$, or 0.8035 gram. From the above illustrations the use of subscripts to the symbols of the elements is perhaps already sufficiently evident, but two further illustrations will here be given. The symbol for cane sugar is $C_{12}H_{22}O_{11}$, which shows that this compound is composed of 12×12 or 144 parts of carbon by weight, to every 1×22 or 22 parts of hydrogen by weight, to every 11×16 or 176 parts of oxygen by weight; in other words every 342 grams of sugar contain 144 grams of carbon plus 22 grams of hydrogen plus 176 grams of oxygen. The symbol for phosphoric acid anhydride, also called phosphorus pentoxide, is P_2O_5. This formula shows that this compound is composed of phosphorus and oxygen in the proportions of 2×31 or 62 grams of the former to 5×16 or 80 grams of the latter.

The use of chemical symbols, then, serves to express quite a number of facts about a chemical compound in compact form. A few of the chemical changes already studied will now be expressed, using chemical symbols and equations.

These will serve to illustrate further how the formulas are used.

(1) When iron filings and sulphur are heated together, ferrous sulphide is formed. Thus:

Fe	+	S	=	FeS
iron		sulphur		ferrous sulphide
56 grams		32 grams		88 grams

(2) When carbon is burned in oxygen or in the air, carbon dioxide is formed. Thus:

C	+	O_2	=	CO_2
carbon		oxygen		carbon dioxide
12 grams		32 grams		44 grams

(3) When phosphorus is burned in oxygen, phosphorus pentoxide is formed. Thus:

P_2	+	O_5	=	P_2O_5
phosphorus		oxygen		phosphorus pentoxide
62 grams		80 grams		142 grams

(4) When sulphuric acid acts on zinc, hydrogen and zinc sulphate are formed. Thus:

H_2SO_4	+	Zn	=	H_2	+	$ZnSO_4$
sulphuric acid		zinc		hydrogen		zinc sulphate
98 grams		65 grams		2 grams		161 grams

(5) When hydrochloric acid is neutralized by sodium hydroxide, sodium chloride and water are formed. Thus:

HCl	+	NaOH	=	NaCl	+	H_2O
hydrochloric acid		sodium hydroxide		sodium chloride		water
36.5 grams		40 grams		58.5 grams		18 grams

(6) When potassium chlorate is heated, it yields oxygen and potassium chloride. Thus:

$KClO_3$	=	3 O	+	KCl
potassium chlorate		oxygen		potassium chloride
122.5 grams		48 grams		74.5 grams

There are really only three different types of chemical change possible, namely:

(1) When two or more substances unite to form one new substance, as in reactions (1), (2), and (3) above. These are all examples of **synthesis.**

(2) When two or more substances are formed by decomposition of a single substance, as in reaction (6) above. This may be termed **analysis,** as contrasted with synthesis.

(3) When two or more substances react with one another to form two or more new substances, as in reactions (4) and (5). These are also termed cases of **double decomposition.**

Chemical symbols, then, are a great aid in that they indicate the chemical composition of a compound at a glance. It must, of course, be kept in mind that *such symbols are only the expression of facts found out by experiment,* and so *the formula of a compound cannot be written till after its composition has been actually ascertained by chemical analysis.*

The formation of salts really appears much simpler when expressed by means of equations in which chemical symbols are used, thus:

$$\underset{\substack{\text{potassium hydroxide}\\ \text{or potash lye}}}{KOH} + \underset{\text{nitric acid}}{HNO_3} = \underset{\substack{\text{potassium nitrate}\\ \text{or saltpeter}}}{KNO_3} + \underset{\text{water}}{H_2O}$$

$$\underset{\substack{\text{sodium hydroxide}\\ \text{or soda lye}}}{2\,NaOH} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{water}}{2\,H_2O}$$

$$\underset{\substack{\text{calcium hydroxide}\\ \text{or slaked lime}}}{Ca(OH)_2} + \underset{\substack{\text{hydrochloric}\\ \text{acid}}}{2\,HCl} = \underset{\text{calcium chloride}}{CaCl_2} + \underset{\text{water}}{2\,H_2O}$$

In regarding the last three equations it will be seen that *an* **acid** *is a compound containing hydrogen which may be replaced by a metal, thus forming a salt.* So potassium nitrate, sodium sulphate, and calcium chloride are typical

salts. Now the hydroxides of potassium, sodium, and calcium are typical *bases.* *A* **base** *is a compound (usually an hydroxide of a metal) which upon reacting with an acid forms a salt and water.* And finally *a* **salt** *is a compound formed when a base acts upon an acid.* We may consider salts as the products formed when the hydrogen of an acid is replaced by a metal or some group of elements that may play the rôle of a metal so far as the process of a salt formation is concerned.

Common salt, NaCl, may be formed by direct union of sodium and chlorine, thus:

$$Na + Cl = NaCl$$

or by neutralization of the base, sodium hydroxide, NaOH, by hydrochloric acid, thus:

$$NaOH + HCl = NaCl + H_2O$$

Calcium sulphate, $CaSO_4$, may be formed by the union of lime, CaO, with sulphuric anhydride, SO_3, thus:

$$CaO + SO_3 = CaSO_4$$

or by the action of slaked lime on sulphuric acid, thus:

$$Ca(OH)_2 + H_2SO_4 = CaSO_4 + 2\ H_2O$$

The metals, then, are in general base-forming substances, and *the non-metals are acid-forming substances, though some metals may at times act as acid-forming elements, and some of the non-metals may exhibit basic properties.*

Among the most common salts are chlorides, sulphates, nitrates, carbonates, phosphates, and silicates. These may be considered as derived respectively from the following acids: hydrochloric acid HCl, sulphuric acid H_2SO_4, nitric acid HNO_3, carbonic acid H_2CO_3, phosphoric acid H_3PO_4,

and silicic acid H_2SiO_3. So, for example, we have sodium chloride NaCl, calcium chloride $CaCl_2$, potassium chloride KCl, magnesium chloride $MgCl_2$, ferric chloride $FeCl_3$, cupric chloride $CuCl_2$, etc., all of which may be considered as derived from hydrochloric acid HCl, the hydrogen of which is replaced by the respective metals. Again, we have sodium sulphate Na_2SO_4, potassium sulphate K_2SO_4, ammonium sulphate $(NH_4)_2SO_4$, copper sulphate $CuSO_4$, calcium sulphate $CaSO_4$, magnesium sulphate $MgSO_4$, ferrous sulphate (also commonly called copperas) $FeSO_4$, etc. These sulphates may all be considered as derived from sulphuric acid H_2SO_4, whose hydrogen has been replaced by the metals or the ammonium group respectively. Similarly we have potassium nitrate KNO_3, sodium nitrate $NaNO_3$, calcium nitrate $Ca(NO_3)_2$, copper nitrate $Cu(NO_3)_2$, ammonium nitrate NH_4NO_3, etc., all of which may be regarded as derived from nitric acid HNO_3. The carbonates like sodium carbonate Na_2CO_3, potassium carbonate K_2CO_3, ammonium carbonate $(NH)_2CO_3$, calcium carbonate $CaCO_3$, lead carbonate $PbCO_3$, etc., may be considered as derivatives of carbonic acid H_2CO_3. The phosphates like calcium phosphate $Ca_3(PO_4)_2$, sodium phosphate Na_2HPO_4, potassium phosphate K_2HPO_4, ferric phosphate $FePO_4$, may all be regarded as derived from phosphoric acid H_3PO_4; whereas silicates like sodium silicate Na_2SiO_3, calcium silicate $CaSiO_3$, etc., may be considered as derivatives of silicic acid. In like manner acetates are derived from acetic acid, arsenates from arsenic acid, lactates from lactic acid, tartrates from tartaric acid, oleates from oleic acid, borates from boric acid, and so on. *Each acid is, then, capable of forming a series of salts which result when the hydrogen of the acid is replaced by the various metals or radicals* (i.e. *groups of elements like ammonium,* NH_4) *that may act as metals.* In turn

each metal may form a series of salts as that metal replaces the hydrogen of the various acids. So, for instance, we have the chloride of copper, the nitrate of copper, the sulphate of copper, the borate of copper, the oleate of copper, the phosphate of copper, the silicate of copper, etc. Each metal may in general form a similar long list of salts with the various acids.

QUESTIONS

1. Define the terms acid, base, salt, and give an example of each.

2. Mention six acids found in nature, stating where they occur.

3. What is an alkali? Give four examples.

4. Write the equation expressing the neutralization of hydrochloric acid by caustic soda.

5. Explain the meaning of the following symbols: H, O, Co, P, Cl, N, K, Hg, Au, Pt, S.

6. What is the law of definite proportions?

7. What is the meaning of the following formulas: H_2O, NaCl, P_2O_5, $C_{12}H_{22}O_{11}$?

8. How much ferrous sulphide, FeS, may be formed from 200 grams of sulphur? How many pounds would this be?

9. How much common salt, NaCl, would be obtained when 2 pounds of sodium are burned in chlorine?

10. How much lime, CaO, would be required to produce 1000 grams of chalk, $CaCO_3$?

CHAPTER VI

THE HALOGENS

THE elements *fluorine, chlorine, bromine, and iodine are called the* **halogens.** They never occur in nature in the uncombined state, but they are frequently met in compounds.

Fluorine is found chiefly as calcium fluoride, *fluor spar*, CaF_2, but it is after all fairly widely distributed in minute amounts in granitic rocks, being a minor constituent of the mineral *apatite*, which consists essentially of calcium phosphate plus a small percentage of calcium fluoride. Fluorine is present in extremely small quantities in soils, from which it gets into plants, and so into animals. In the enamel of the teeth of the latter notable amounts of fluorine are always present. By treating calcium fluoride with sulphuric acid hydrogen fluoride, **hydrofluoric acid**, HF, is formed, thus:

$$\underset{\text{calcium fluoride}}{CaF_2} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{calcium sulphate}}{CaSO_4} + \underset{\text{hydrofluoric acid}}{2\,HF}$$

This acid is a gas which readily dissolves in water. It attacks glass, sand, and silicates in general, forming silicon fluoride, which is volatile, and metallic fluorides with any bases that may be present. Thus its action on calcium silicate is shown by the following equation:

$$\underset{\text{calcium silicate}}{CaSiO_3} + \underset{\text{hydrofluoric acid}}{6\,HF} = \underset{\text{calcium fluoride}}{CaF_2} + \underset{\text{silicon tetrafluoride}}{SiF_4} + \underset{\text{water}}{3\,H_2O}$$

Hydrofluoric acid is poisonous. The gas and also the aqueous solution of the latter are used for *etching glass*, — diamond ink. Hydrofluoric acid is also used in the laboratory for the purpose of decomposing silicates in chemical analysis. Fluorine itself may be obtained by passing the electric current through a solution of potassium fluoride and hydrofluoric acid in water. Fluorine is a light-colored greenish yellow gas of extremely pungent odor. It unites with all elements directly, except with oxygen. It is probably the most active of all of the chemical elements.

Chlorine *is found mainly in common salt*, sodium chloride, NaCl, the occurrence of which will be discussed in connection with sodium. By treating common salt with sulphuric acid, **hydrochloric acid** is formed, thus:

$NaCl$	$+$	H_2SO_4	$=$	$NaHSO_4$	$+$	HCl
sodium chloride		sulphuric acid		sodium acid sulphate		hydrochloric acid

Hydrochloric acid, also called **muriatic acid** or spirit of salt, is a colorless pungent gas, which fumes in the air, for it has a strong affinity for water and condenses the latter from the air to drops. Thus the fumes are really minute droplets of a solution of hydrochloric acid gas in water. At ordinary temperatures and atmospheric pressure, one volume of water will dissolve about four hundred volumes of hydrochloric acid gas. *Hydrochloric acid is a powerful acid.* With bases it reacts, forming chlorides and water. In the human stomach there is found a 0.33 per cent solution of hydrochloric acid, which aids in the digestion of food. Hydrochloric acid is made in large quantities on a commercial scale in connection with the LeBlanc soda process (which see). The saturated aqueous solution has the specific gravity 1.19 and contains 38 per cent of pure hydrochloric acid by weight.

Chlorine is prepared by abstracting hydrogen from hydrochloric acid. This may be done by simply passing the electric current through an aqueous hydrochloric acid solu-

FIG. 12. — A carboy of hydrochloric acid.

tion, whereupon chlorine is evolved on one pole and hydrogen on the other. The common way, however, is to oxidize hydrochloric acid by means of a suitable oxidizing agent. As such manganese dioxide is usually chosen, thus:

$$\underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{hydrochloric acid}}{4\,HCl} = \underset{\text{manganous chloride}}{MnCl_2} + \underset{\text{water}}{H_2O} + \underset{\text{chlorine}}{Cl_2}$$

The manganese dioxide and hydrochloric acid are gently heated together in a flask or retort. Chlorine is a greenish yellow gas with a very pungent odor. It is extremely active chemically, forming chlorides with many of the elements by direct union. So with sodium it forms sodium chloride, $NaCl$; with copper, cupric chloride, $CuCl_2$; with phosphorus, phosphorus chloride, PCl_3; with sulphur, sulphur chloride, S_2Cl_2; etc. On water chlorine reacts in the sunlight, forming oxygen and hydrochloric acid, thus:

$$H_2O + Cl_2 = 2\,HCl + O$$

The oxygen thus set free oxidizes many substances, destroying colors, microörganisms, etc. For this reason **chlorine water** (a solution of chlorine gas in water) is a bleaching agent and an antiseptic. When chlorine gas is conducted into slaked lime, so-called **chloride of lime, bleaching powder,** is formed, thus:

$$\underset{\text{slaked lime}}{Ca(OH)_2} + \underset{\text{chlorine}}{Cl_2} = \underset{\substack{\text{chloride of lime}\\ \text{bleaching powder}}}{CaCl(OCl)} + \underset{\text{water}}{H_2O}$$

Bleaching powder is a powerful antiseptic and bleaching agent. Its action depends upon the fact that it will readily yield its oxygen and pass over into calcium chloride, $CaCl_2$, thus:

$$CaCl\,.\,OCl = CaCl_2 + O$$

The oxygen destroys the color of many dyestuffs, and also kills microörganisms. Bleaching powder is manufactured on a large scale commercially, and is used in bleaching fabrics, paper pulp, etc. It is also used to rid the soil of undesirable organisms.

Bromine is mainly found as sodium bromide, $NaBr$, in connection with sodium chloride. It may be prepared by methods that are entirely similar to those used in making

chlorine. Thus hydrobromic acid is formed when sulphuric acid acts on sodium bromide:

$$\underset{\text{sodium bromide}}{NaBr} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{acid sodium sulphate}}{NaHSO_4} + \underset{\text{hydrobromic acid}}{HBr}$$

By treating manganese dioxide with hydrobromic acid, bromine is obtained:

$$\underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{hydrobromic acid}}{4\,HBr} = \underset{\text{water}}{2\,H_2O} + \underset{\text{manganous bromide}}{MnBr_2} + \underset{\text{bromine}}{Br_2}$$

Bromine forms bromides by direct union with many of the elements. Potassium bromide, KBr, is used in medicine and photography. Bromine itself is a brownish liquid of sp. gr. 3.188 at 0°. It boils at 59° C. About 250,000 pounds of it are produced yearly in the United States. It is used as an antiseptic, also in making aniline dyes, bromides, etc. **Hydrobromic acid** when pure is a colorless gas soluble in water. Its properties are similar to those of hydrochloric acid, but it is weaker than the latter. *Bromine is set free from bromides by chlorine,* thus:

$$\underset{\text{sodium bromide}}{NaBr} + \underset{\text{chlorine}}{Cl} = \underset{\text{sodium chloride}}{NaCl} + \underset{\text{bromine}}{Br}$$

Bromine turns starch paste yellow.

Iodine is found in the ashes of sea weeds, also as sodium iodate, $NaIO_3$, in connection with Chili saltpeter. Iodine is a crystalline, grayish black solid having almost a metallic luster. It dissolves sparingly in water, but copiously in alcohol. The alcoholic solution is called tincture of iodine. It is used in medicine as an antiseptic and counter-irritant. With the metals and some of the other elements iodine forms iodides. Thus we have sodium iodide NaI, potassium iodide KI, silver iodide AgI, calcium iodide CaI_2, phosphorus

iodide PI_3, etc. From sodium iodide, iodine may readily be obtained by treatment with sulphuric acid. Thus hydrogen iodide, **hydriodic acid**, HI, is formed, which, however, at once attacks the sulphuric acid present, forming sulphorous acid, H_2SO_3, and water, thus:

$$\underset{\text{sodium iodide}}{NaI} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{acid sodium sulphate}}{NaHSO_4} + \underset{\text{hydriodic acid}}{HI}$$

and

$$\underset{\text{sulphuric acid}}{H_2SO_4} + \underset{\text{hydriodic acid}}{2\,HI} = \underset{\text{sulphurous acid}}{H_2SO_3} + \underset{\text{water}}{H_2O} + \underset{\text{iodine}}{2\,I}$$

Iodine turns starch paste blue. This fact is used in testing for starch. It is evident that one may also use starch paste in testing for iodine. The vapors of iodine are of a beautiful violet color, from which fact the element has received its name.

The thyroid gland contains iodine in the form of a complex compound, thyroiodine. In the treatment of goiter and other diseases connected with the thyroid gland, iodine compounds, like potassium iodide, or extract of the thyroid gland of the sheep, are frequently prescribed by the physician.

Iodine may be liberated from iodides by treating the latter with bromine or chlorine, thus:

$$\underset{\text{potassium iodide}}{KI} + \underset{\text{chlorine}}{Cl} = \underset{\text{potassium chloride}}{KCl} + \underset{\text{iodine}}{I}$$

QUESTIONS

1. Name the halogens, and one important compound of each found in nature.

2. (*a*) How may hydrofluoric acid be prepared?
   (*b*) What use is made of this compound?

3. Describe hydrochloric acid and tell how it is prepared.

4. What is hydrochloric acid used for? What other names does this compound have?

5. How may chlorine be made? Describe this element.

6. Explain how chlorine bleaches.

7. What is bleaching powder? Upon what fact does its action depend?

8. What use is made of bromine? How may it be prepared?

9. Compare the properties of iodine with those of the other elements.

10. Mention some of the uses of iodine. Compute how much iodine would be obtained when 100 grams of chlorine act upon an excess of potassium iodide.

CHAPTER VII

SULPHUR, PHOSPHORUS, ARSENIC, ANTIMONY, AND BISMUTH

Sulphur is found in nature in the free or uncombined state in large quantities. The largest deposits occur in Sicily, Texas, and Louisiana. Sulphur was well known even in ancient times. It frequently occurs around the brims of craters of volcanoes, and so is also called **brimstone.** In the market it may be obtained in sticks as *roll sulphur,* and also in powdered form as "flowers of sulphur." Sulphur melts at 114°.5, forming a lemon-colored liquid, which on further heating turns brown and becomes viscous at about 160° to 200° C. Finally on further heating this liquid again becomes limpid and boils at 450° C., giving off brown vapors. Sulphur will burn in the air or in oxygen, forming **sulphur dioxide,** thus:

$$S + O_2 = SO_2$$

Sulphur dioxide is a gas of suffocating odor. In it living beings cannot exist, and so it is used as a disinfecting agent. As such it is cheap. It is very commonly used in fumigating rooms and houses that have been occupied by persons having contagious diseases. For this purpose it may be had on the market in liquid form in tin cans, Fig. 13. These are soldered shut and have a small short lead tube attached to them, the end of which can be nipped off. The

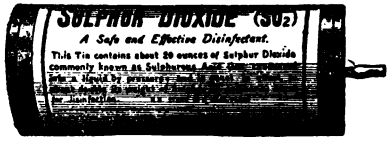


FIG. 13. — A can filled with liquid sulphur dioxide.

gas can then make its escape into the room. The latter should be kept tightly closed for several hours after it has been filled with sulphur dioxide. The gas may also be prepared by burning sulphur in the rooms to be disinfected. Care must be taken in this process not to set the house on fire (see Chapter XXI).

Sulphur dioxide is also used as a bleaching agent. It is especially used for bleaching straw, feathers, and fabrics

FIG. 14. — Pumping sulphur from a Louisiana sulphur well.

that would be injured by bleaching powder. It used to be employed also in bleaching food products, such as dried apples, and grains discolored by fire, for example; but this is not to be recommended on account of the poisonous nature of the substance. *Sulphur dioxide bleaches by uniting chemically with the substance whose color is altered.* The action is thus quite different from that of bleaching with either hydrogen peroxide or bleaching powder, which bleach because of oxidation of the coloring matter.

Sulphur is also used in making fireworks, black gunpowder, both soft and hard rubber, carbon bisulphide, lime sulphur mixtures for spraying shrubs and trees, and **sulphuric acid.** Thousands of tons of the latter compound are produced annually. In its production sulphur is first converted into sulphur dioxide by burning it in the air. The sulphur dioxide is then mixed with air and this mixture is passed over finely divided platinum heated from 400° to 450° C. In *contact* with this hot, finely divided platinum the sulphur dioxide unites further with the oxygen of the air, forming **sulphur trioxide**, thus:

$$SO_2 + O = SO_3$$

Sulphur trioxide forms long, white, crystalline fibers, which melt at 14°.8 C. It takes on water very greedily, forming sulphuric acid, thus:

$$SO_3 + H_2O = H_2SO_4$$

Sulphuric acid, also called **oil of vitriol**, is a heavy viscous liquid, being 1.838 times as heavy as water at 15° C. It has great affinity for water and will dissolve in the same with liberation of much heat. *Care must consequently be used in pouring sulphuric acid into water. Always pour the acid*

gradually *into the water (never the water into the acid) and stir well after each addition of acid.* If the water is poured into the acid, or if too much of the acid is poured into the water at a time, the heat evolved is so great that liquid is liable to be thrown out of the container and injure the person. Sulphuric acid is a very important article of commerce. With most of the metals, ammonia, and other basic substances, it forms salts called sulphates. Thus we have copper sulphate $CuSO_4$, calcium sulphate $CaSO_4$, barium sulphate $BaSO_4$, ammonium sulphate $(NH_4)_2SO_4$, etc. *Besides its ability to unite with basic substances to form* **sulphates,** *the most striking property of sulphuric acid is its attraction for water,* which has already been mentioned. So great is the affinity of sulphuric acid for water that it will abstract the latter from wood, sugar, starch, meat, and in fact from all other animal and vegetable substances, thus charring them. *Sulphuric acid is further used in making superphosphate fertilizers, in preparing hydrochloric acid and soda by the LeBlanc soda process, in manufacturing ether, aniline dyes, guncotton, dynamite, smokeless powder, and many other compounds in which a strong drying agent is required.* Upwards of four million tons of sulphuric acid are manufactured annually. Besides the process of making sulphuric acid, which has already been described and which is known as the **contact process,** sulphur dioxide is also oxidized by means of nitric acid and oxides of nitrogen in the presence of water vapor. This is an older process which is carried on in lead-lined chambers, lead being but very slightly attacked by sulphuric acid. The **lead chamber process** is still used to a large extent, especially for making sulphuric acid of 60 to 78 per cent strength.

Sulphuric acid will react with many salts, forming sulphates and liberating the acids of such salts, thus:

$$\underset{\text{sodium chloride}}{NaCl} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{acid sodium sulphate}}{NaHSO_4} + \underset{\text{hydrochloric acid}}{HCl}$$

$$\underset{\text{potassium nitrate}}{KNO_3} + H_2SO_4 = \underset{\text{potassium acid sulphate}}{KHSO_4} + \underset{\text{nitric acid}}{HNO_3}$$

Fuming sulphuric acid, also called *Nordhausen sulphuric acid,* is sulphuric acid in which additional sulphur trioxide has been dissolved.

Sulphurous acid is formed when sulphur dioxide dissolves in water, thus: $SO_2 + H_2O = H_2SO_3$

From it salts called **sulphites** are derived. Thus we have sodium sulphite Na_2SO_3, sodium acid sulphite, also called sodium bisulphite, $NaHSO_3$, calcium sulphite $CaSO_3$, etc. *The sulphites are used in photography and in the preparation of wood pulp in the paper industry.* Sulphites also used to be employed as food preservatives, but this is now forbidden by law, for they are injurious to health.

Hydrogen sulphide is a compound consisting of hydrogen and sulphur. Its composition is expressed by the formula H_2S. This gas is formed in rotten eggs, in manure pits, and wherever animal and vegetable matter is decaying without proper access of air. Hydrogen sulphide is poisonous. It dissolves in water, one volume of the latter absorbing about three volumes of the gas. Hydrogen sulphide is 1.19 times as heavy as air. It will burn, forming water and sulphur dioxide, thus:

$$H_2S + 3\,O = H_2O + SO_2$$

Hydrogen sulphide may be prepared by treating ferrous sulphide with hydrochloric or sulphuric acid, thus:

$$FeS + 2\,HCl = FeCl_2 + H_2S$$
$$FeS + H_2SO_4 = FeSO_4 + H_2S$$

When introduced into solutions of salts of the heavy metals, hydrogen sulphide forms characteristic precipitates consisting of the sulphides of the metals, and it is consequently much used in testing for heavy metals in chemical analysis, thus:

$$CuSO_4 + H_2S = \underset{\text{copper sulphide}}{CuS} + H_2SO_4$$

$$2\,AsCl_3 + 3\,H_2S = \underset{\text{arsenious sulphide}}{As_2S_3} + 6\,HCl$$

Carbon bisulphide is a colorless liquid of specific gravity 1.26. It boils at 47° and is *very inflammable. Its vapors mixed with air are explosive, and consequently no flame or spark must be present when carbon bisulphide is being used.* It will not dissolve in water. It is used as a solvent for fats and oils, also for the extermination of ants and other insect pests. Carbon bisulphide is prepared by heating carbon and sulphur together out of contact of the air, usually in an electric furnace. The gas formed is then condensed by means of cold water.

Sulphur occurs in nature also as calcium sulphate $CaSO_4$, **gypsum** $CaSO_4 . 2\,H_2O$, and also in combination with metals as sulphides like **iron pyrites** or **fool's gold** FeS_2, sulphide of lead or **galenite** PbS, zinc sulphide or **black jack** ZnS, etc. From pyrites sulphur dioxide is obtained by heating the mineral in the air. Much of the sulphur dioxide used in the manufacture of sulphuric acid is obtained from this source. Nearly two and one half million tons of gypsum are produced annually in the United States alone. *It is used as land plaster, as wall plaster, stucco,* etc.

Finally *all plants and animals contain sulphur. Without it they could not live.* Sulphur is a constituent of albumen. Thus muscular tissues, horns, hoofs, hair, nerves, eggs, and seeds contain sulphur. In urine sulphur is present in the

form of sulphates, for as the animal lives, the sulphur compounds in its tissues are continually oxidized to sulphates, which are then eliminated by means of the kidneys. Some plants, like onions, garlic, mustard, skunk cabbages, radishes, contain odoriferous sulphur compounds. In order to thrive well they require a soil which contains an adequate amount

FIG. 15. — A group of sulphur crystals on limestone.

of sulphur. When plants and animals decay, the sulphur compounds they contain gradually decompose, forming hydrogen sulphide when the sulphur is relatively abundant and access to the air is limited, and sulphates when there is proper access of air. When these sulphates are again put into the soil, as in manuring or in treating with land plaster (*i.e.* gypsum), the rootlets of plants again take them up, and transform them into the various sulphur

compounds of their tissues, and thus the **sulphur cycle** is completed.

Sulphur springs contain hydrogen sulphide, which is easily detected by its foul odor. Many mineral waters contain calcium sulphate, some contain sodium sulphate, Na_2SO_4, others like those at Epsom contain magnesium sulphate, $MgSO_4$. The latter are bitter to the taste and have a laxative effect. *Magnesium sulphate is called Epsom salt, and is sometimes taken as a purgative.*

Phosphorus *never occurs in nature in the free state.* It is, however, quite widely distributed in small quantities in compounds. It occurs as **calcium phosphate,** $Ca_3(PO_4)_2$, in **phosphate rock** in the Carolinas and some of the Western states. Small amounts of phosphates occur also in all fertile soils, in iron ores (from which it passes into blast furnace slags when the ores are reduced), and in many granitic rocks, clays, etc. *Bones consist of calcium phosphate to the extent of about 80 per cent. Phosphorus is an important and indispensable part of the tissues of all plants and animals.* In the nerves, brain, blood, muscles, and in fact in all animal parts that are concerned in locomotion or reproduction, phosphorus is found. In seeds phosphorus is always present, particularly in the embryos, and *without phosphorus plants cannot grow,* hence

FIG. 16. — A piece of phosphate rock.

its importance in the soil. In the bodies of plants and animals phosphorus is combined with carbon, hydrogen, oxygen, nitrogen, and sulphur. The nerves and the brain are especially rich in a complex compound, **lecithine**, whose composition is expressed by the formula $C_{44}H_{90}NPO_9$.

FIG. 17. — Barley growing with and without proper food.

(1) Complete manure. (2) No nitrogen. (3) No phosphorus.
(4) No potassium. (5) No calcium. (6) No magnesium.

As animals live these complex compounds are gradually oxidized, phosphates being formed, which are then eliminated by the kidneys. Thus urine always contains phosphates, especially calcium and potassium phosphates. When these are then returned to the land as fertilizer, they are again taken up by the roots of plants and converted into complex phosphorus compounds in their tissues and seeds. These in turn are eaten by animals, and so the **phosphorus cycle** completes itself. It is clear, however, that considerable amounts of phosphorus are never returned to the land, being washed away continually together with other valuable fertilizer material into the waterways, which finally empty into the ocean. *The sewage of our large cities in particular represents a colossal waste. In order to keep up the supply of phosphorus, bones, blast furnace slags, and ground phosphate rock are employed as fertilizers.* When bones are treated with sulphuric acid, so-called **superphosphate** is produced, thus:

$$\underset{\text{calcium phosphate}}{Ca_3(PO_4)_2} + \underset{\text{sulphuric acid}}{2\,H_2SO_4} = 2\,CaSO_4 + \underset{\text{superphosphate}}{CaH_4(PO_4)_2}$$

This is more soluble in the waters of the soil and hence more available to plants.

Phosphorus itself is produced by heating calcium phosphate with silica (sand) and carbon. These are finely ground and intimately mixed and then heated out of contact with the air in earthenware retorts or in a suitable electric furnace. The following change takes place:

$$\underset{\text{calcium phosphate}}{Ca_3(PO_4)_2} + \underset{\text{silica}}{3\,SiO_2} + \underset{\text{carbon}}{5\,C} = \underset{\text{calcium silicate}}{3\,CaSiO_3} + \underset{\text{carbon monoxide}}{5\,CO} + \underset{\text{phosphorus}}{2\,P}$$

Phosphorus is a yellowish, translucent, waxlike solid. Under water it may be melted at 44° C. It catches fire in the air at about 35° C., and hence it is always kept and trans-

ported out of contact with the air, usually under water in air-tight cans. *Phosphorus is very poisonous*, 0.1 gram being sufficient to produce death in the case of an adult. *Phosphorus burns are dangerous*, and they heal very slowly. Phosphorus should consequently be handled with great care. *Forceps should always be used when handling phosphorus, and small amounts, usually not larger than a pea, should be employed in experiments by the beginner.*

Phosphorus is used to poison rats and other vermin, and in making **matches**. For the latter purpose yellow phosphorus should not be employed, for the workers in match factories are frequently poisoned by working with it, the poisoning being evidenced by an enlargement of the liver and a gradual destruction of the jawbones. When heated out of contact with the air to from 250° to 300° C., yellow phosphorus changes to a red powder, called **red phosphorus**. This allotropic form of phosphorus is far less dangerous than the yellow variety. It will not take fire in the air till heated to about 200° C., and moreover it is not poisonous. It is frequently employed in the chemical laboratory in preparing phosphorus compounds, and it is also used in making **safety matches**. The latter commonly contain a mixture of potassium chlorate, potassium bichromate, powdered glass, and glue or dextrine, as the match head, whereas on the friction surface on the box there is a mixture of red phosphorus, antimony sulphide, manganese dioxide, and glue. The powdered glass increases the friction, which raises the temperature to the point at which the match head bursts into flame. *Safety matches ignite only by rubbing on the specially prepared surface on the box.* They were first prepared in Sweden and so are often called *Swedish matches*. *Matches containing yellow phosphorus in the head will ignite by rubbing on any surface.* They are dangerous because they may cause

conflagrations, and also because of their poisonous nature many lives are sacrificed in manufacturing them. Phosphorus matches first came into use in 1832, and now some 3000 tons of phosphorus are annually used up in making matches.

Phosphoric acid *is produced by treating calcium phosphate with sulphuric acid,* thus:

$$\underset{\text{calcium phosphate}}{Ca_3(PO_4)_2} + \underset{\text{sulphuric acid}}{3\ H_2SO_4} = \underset{\text{calcium sulphate}}{3\ CaSO_4} + \underset{\text{phosphoric acid}}{2\ H_3PO_4}$$

It is a sirupy liquid of specific gravity 1.88. It is a pronounced acid and forms phosphates with the metals, with ammonia, and other bases. On heating phosphoric acid it loses water and forms metaphosphoric acid, HPO_3, thus:

$$H_3PO_4 = HPO_3 + H_2O$$

Metaphosphoric acid is analogous to nitric acid, HNO_3.

When phosphorus is burned in the air or in oxygen, phosphorus pentoxide P_2O_5 is formed, thus:

$$2\ P + 5\ O = P_2O_5$$

The latter on treatment with water forms phosphoric acid, thus:

$$P_2O_5 + 3\ H_2O = 2\ H_3PO_4$$

As phosphorus is a valuable constituent of fertilizers, the percentage content in the latter should always be considered in making purchases. The per cent of phosphorus is sometimes stated as per cent P, which means the element itself, but more frequently it is stated on the label as per cent phosphorus pentoxide, P_2O_5, which is erroneously called phosphoric acid. It is to be remembered that P_2O_5 contains only 62 pounds of phosphorus to 80 pounds of oxygen. That is to say, it contains only 62 pounds of phosphorus in

every 142 pounds. Dealers like to express the phosphorus content of fertilizers as per cent of P_2O_5, for, of course, the figure appears much larger than when expressed as per cent P.

Arsenic is a steel-gray to black, brittle substance having a metallic luster. It burns in the air emitting a garlic-like odor and thus forms a white powder, **arsenious oxide**, or **white arsenic**, As_2O_3, which floats in the air as a white cloud. Indeed, *white arsenic is the most important compound of arsenic.* It is quite common on the market and relatively inexpensive. Like all compounds of arsenic, it is very poisonous. White arsenic, arsenious acid, or " arsenic," as it is also often called in commerce, is used as rat poison. About 1500 tons are produced yearly in the United States. It is slightly soluble in water and has a sweetish, sickening taste. *The antidote for arsenic poisoning is freshly precipitated ferric hydroxide.* This forms an insoluble mass with arsenious oxide.

Arsenic compounds are used almost entirely as poisons for vermin and insect pests. Two compounds of arsenic are at present very commonly employed for this purpose, namely, **Paris green**, whose composition is $Cu_3As_2O_6 . Cu(C_2H_3O_2)_2$, being a double salt of cupric arsenite and cupric acetate, and **lead arsenate** $Pb_3(AsO_4)_2$. Both Paris green and lead arsenate are but sparingly soluble in water. They are used in exterminating potato bugs and various other insects. Directions for preparing the proper mixtures for such spraying will be given later. *It should be kept in mind that both compounds are strong poisons and must be used with proper care.*

Antimony is a brittle metal. It is used in making **type metal** and **Babbitt metal** for machine bearings. As molten antimony solidifies, it expands and fills the molds, hence it is valuable in many alloys; moreover, it is hard and when

mixed with lead and tin forms an alloy which is particularly suitable for type, the composition being 25 per cent antimony, 50 per cent lead, and 25 per cent tin. Babbitt metal contains 70 to 90 per cent lead, alloyed with tin and antimony. **Bullets** and **shot** consist of lead alloyed with from 0.2 to 0.4 per cent arsenic.

Bismuth, too, is a brittle metal. It is used in preparing low-melting alloys which are employed in making readily fusible plugs in the pipes of automatic sprinklers for fire protection. So, for instance, Wood's metal, consisting of 1 part tin, 2 parts lead, 1 part cadmium, and 4 parts bismuth, melts at 60°.5 C. Bismuth compounds are used in medicine, the most common preparation being **bismuth subnitrate**, $BiO . NO_3 . BiO . OH$, a white powder which is difficultly soluble in water and practically tasteless. It is prescribed in cases of dysentery and other disturbances of the alimentary canal. It is also sometimes used as a face powder.

QUESTIONS

1. How does sulphur occur in nature?
2. Describe how this element behaves on heating.
3. What is formed when sulphur is burned in the air? Describe the properties and uses of this substance.
4. What use is made of sulphur?
5. What are the two methods by which sulphuric acid is manufactured on a large scale?
6. Describe the properties of sulphuric acid. What use is made of this substance?
7. What are sulphites? Sulphates? Give an example of each.
8. What is hydrogen sulphide? How is it prepared?
9. How is carbon bisulphide made? What is it used for?
10. Discuss the rôle of sulphur in plants and animals.

11. In what forms is phosphorus found in nature?

12. Discuss the rôle of phosphorus in plant and animal life.

13. What use is made of phosphorus in the arts?

14. What is superphosphate fertilizer?

15. How much phosphoric acid, H_3PO_4, could be produced from one ton of calcium phosphate, $Ca_3(PO_4)_2$, the essential compound in phosphate rock?

16. How much phosphorus is there in 1 pound of phosphorus pentoxide, P_2O_5? How much phosphoric acid, H_3PO_4, could be made from 1 pound of P_2O_5?

17. What is white arsenic?

18. In what other commercial compounds does arsenic occur? What are these used for?

19. How much arsenic is there in a pound of Paris green?

20. What are antimony and bismuth, and what use is made of them?

CHAPTER VIII

BORON AND SILICON

Boron occurs in nature in boric acid and its sodium and calcium salts. It is never found in the uncombined state. The element boron is a brown, brittle solid. It may be dissolved in molten aluminum, and when the mass cools boron separates out in form of crystals that are almost as hard as diamond. **Boric acid** has the composition H_3BO_3. It may be prepared from **borax**, $Na_2B_4O_7 . 10\ H_2O$, by treating a hot concentrated solution of the latter with either hydrochloric or sulphuric acid. The boric acid separates out in the form of crystalline flakes on cooling, thus:

$$\underset{\text{sodium tretraborate}}{Na_2B_4O_7} + \underset{\text{hydrochloric acid}}{2\ HCl} + \underset{\text{water}}{5\ H_2O} = \underset{\text{sodium chloride}}{2\ NaCl} + \underset{\text{boric acid}}{4\ H_3BO_3}$$

Boric acid and borax are by far the most important compounds of boron. Boric acid is known also as **boracic acid.** It dissolves in water. At room temperature the saturated solution contains about 4 per cent. It is frequently used as an antiseptic wash, especially in treating the eyes, for it is a very mild agent. *Compresses of saturated boric acid solution are frequently applied on parts of the body that are swollen because of blood poisoning.* Powdered boric acid feels slippery to the touch, and consequently is at times used to spread on the floors of dance halls. It has also been used to preserve milk, meats, fish, etc.; but as it is injurious to health, the practice is now forbidden by law. Boric acid has prac-

tically no effect on litmus, for it is an extremely weak acid; it turns turmeric paper reddish brown, however, and this serves as a delicate test for boric acid. When the turmeric paper thus reddened is afterward treated with caustic alkali solution, it turns to a dark greenish black.

Borax is alkaline toward litmus, for it is a salt of a very weak acid and a powerful base. The salt forms beautiful crystals, which crumble to powder on standing in the air, because they lose water. Over forty thousand tons of borax are produced in the United States annually. The chief deposits of colemanite, $Ca_2B_2O_{11} \cdot 5\ H_2O$, from which much borax is made, occur in California and Oregon.

FIG. 18. — A group of borax crystals.

In the laundry, borax is used for softening water and for enhancing the gloss of starch in ironing. It is also used as an antiseptic, as a mordant in dyeing fabrics, and as a flux in welding iron and brazing metals together. Like boric acid, it has also been used for preserving foods, but this practice is to be strongly condemned. In making glazes and enamels for pottery and enameled iron ware, borax and boric acid are often used, also in the production of certain kinds of hard glass.

While boron compounds after all are found in relatively small quantities on the earth, compounds of silicon are extremely abundant, for they make up by far the larger portion

of the solid part of the earth. The element **silicon** never occurs in the uncombined form. It is a brittle solid having metallic luster. Its specific gravity is 2.49. It is so hard that one can scratch glass with it. It is now made on a large scale by heating silicon dioxide, quartz sand, with coke in the electric furnace, thus:

$$\underset{\text{silicon dioxide}}{SiO_2} + \underset{\text{carbon}}{2\,C} = \underset{\text{carbon monoxide}}{2\,CO} + \underset{\text{silicon}}{Si}$$

At present silicon is used mainly in the manufacture of steel, where it serves as a reducing agent.

FIG. 19. — Molten silicon as it is being tapped from an electric furnace.

Silicon dioxide, SiO_2, also called **silica**, *is the most important compound of silicon. Quartz, quartzite, flint, chalcedony, opal, amethyst, agate, carnelian, sandstone, white sand, especially that on the seashore and the desert, are almost entirely*

silicon dioxide. Silica is hard and brittle, and is much used as an abrasive. **Sandpaper** consists of quartz sand glued on paper. Silica may be melted to a clear glass in the oxyhydrogen flame. Utensils of such glass are still quite costly. They are mainly used in laboratories. Although **quartz glass** is quite brittle, it will not break when exposed to sudden changes of temperature. So a dish of quartz glass may be made red hot and then thrust into cold water without breaking the dish. *Quartz sand is also employed in making glass, porcelain, portland cement, and ordinary lime mortar.*

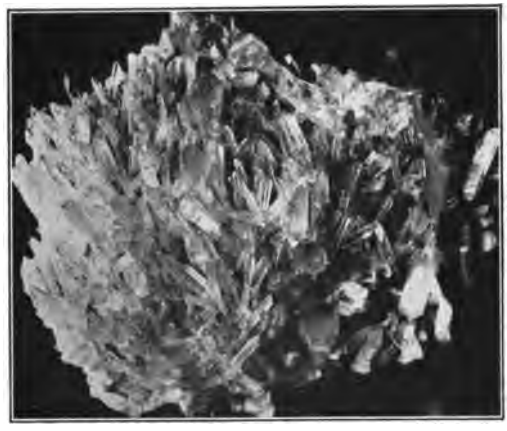


FIG. 20. — Quartz crystals from Hot Springs, Arkansas.

When pulverized quartz sand is fused with soda (sodium carbonate), **sodium silicate** is formed, thus:

$$\underset{\text{silica}}{SiO_2} + \underset{\text{soda}}{Na_2CO_3} = \underset{\substack{\text{sodium silicate}\\ \text{water glass}}}{Na_2SiO_3} + \underset{\text{carbon dioxide}}{CO_2}$$

Sodium silicate is also called **water glass.** *It is readily soluble in water, and its solutions are employed in preserving eggs, and in cementing asbestos fibers and asbestos paper together.* When sodium silicate solution is treated with

hydrochloric or some other strong acid, **silicic acid** is formed, and this separates from concentrated solutions in the form of flakes or jelly, thus:

$$\underset{\text{sodium silicate}}{Na_2SiO_3} + \underset{\text{hydrochloric acid}}{2\,HCl} = \underset{\text{silicic acid}}{H_2SiO_3} + \underset{\text{sodium chloride}}{2\,NaCl}$$

At ordinary temperatures silicic acid is a very weak acid. It does not affect litmus, and has no taste. In water it is but very slightly soluble. However, when the sodium silicate solution used to prepare the silicic acid (see above equation) is only 1 per cent strong, silicic acid will not separate on acidifying with hydrochloric acid. By placing this solution in an animal bladder, like a hog's bladder, and immersing this in a pail of water, the sodium chloride and excess of hydrochloric acid in the solution will pass through the walls of the bladder into the water on the outside of the latter. By renewing the water in the pail constantly, there will finally remain behind in the bladder only silicic acid. Instead of a bladder, a bag made of parchment paper could also be used. This process of thus separating the silicic acid from the other ingredients in the solution was first used by Thomas Graham. It is called **dialysis.** By carefully evaporating off some of the water, the silicic acid solution may be concentrated till it contains about 10 per cent silicic acid. However, the latter is very apt to separate in the form of a jelly on standing.

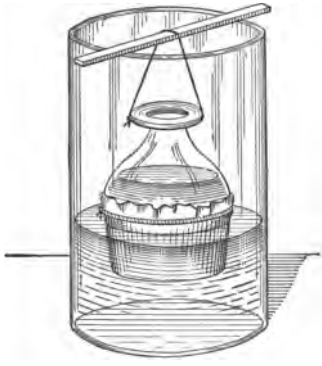


FIG. 21. — Preparing silicic acid by dialysis.

Salts of silicic acid are exceedingly abundant in nature.

The only soluble salts of this acid are those of sodium and potassium and these are prepared artificially. *The naturally occurring silicates are insoluble.* Among some of the most important of these may be mentioned: *potash feldspar* $KAlSi_3O_8$, *soda feldspar*, $NaAlSi_3O_8$, *lime feldspar* $CaAl_2Si_2O_8$, *kaolin or clay* $H_2Al_2Si_2O_8 . 2\,H_2O$, *serpentine* $Mg_3Si_2O_7$, *olivine* Mg_2SiO_4, *meerschaum* $Mg_2Si_3O_8 . 4\,H_2O$, *soapstone or talc* $Mg_3H_2Si_4O_{12}$, *asbestos* $Mg_3Si_2O_7 . 2\,H_2O$, *hornblende* $Mg_2CaFeSi_4O_{12}$, and *mica* $KH_2Al_3Si_3O_{12}$. **Granitic rocks** consist of crystals of quartz, feldspar, and mica. Our fine *clay soils are formed by the disintegration of the granitic rocks,* partly by the action of glaciation in earlier geological ages, partly by the process of weathering, that is, by the gradual action of water, air, and wind, particularly by alternate freezing and thawing of water in the crevices of the rocks. Because of their content of potassium and also of phosphorus, for granitic rocks and the clay that have come from their disintegration always contain minor amounts of calcium phosphate, clay soils are in general quite desirable for agricultural purposes.

While sand and the naturally occurring silicates are all only very slightly soluble in water, yet they are by no means entirely insoluble. So *all terrestrial waters like those of springs, wells, rivers, lakes, and the ocean, as well as all soil waters, contain silicates in solution.* From these dissolved silicates, plants derive in a considerable part the mineral matter which constitutes the ash that remains when they are incinerated. *Silica itself is frequently met in the ash of plants.* It is found particularly in the ash of the stalks of the cereals, of grasses, of bamboo and other canes, to which it gives stability. Often half of the ash consists of silica. In feathers, in the hair of animals, and in the shells of various crustaceans, silica abounds. About 40 per cent of the ash

of the feathers of birds consists of silica. *Deposits of diatomic or* **infusorial earth** *consist of the siliceous remains of minute organisms.* This infusorial earth is used in making **dynamite**, which is essentially nitroglycerine mixed with infusorial earth and then molded into sticks. In the plant and animal bodies silica is doubtless combined with other elements in the form of quite complex compounds. The exact nature of these has not yet been determined.

Hydrofluoric acid will attack silica, forming a volatile compound, silicon tetrafluoride and water, thus:

$$\underset{\text{silica}}{SiO_2} + \underset{\text{hydrofluoric acid}}{4\,HF} = \underset{\text{silicon tetrafluoride}}{SiF_4} + \underset{\text{water}}{2\,H_2O}$$

For this reason hydrofluoric acid is much used in the laboratory in analyzing silicates. It is also employed in etching silica, silicates, glass, and porcelain. The etching depends upon the fact that silicon tetrafluoride is formed, and that fluorides of any bases that may be present are simultaneously produced. The art of thus etching glass has been known for centuries.

QUESTIONS

1. What is borax? What use is made of it?
2. How is boric acid prepared? What is it used for?
3. What is the most abundant compound of silicon? How much silicon do 100 lb. of it contain?
4. Make a list of the most important things in the manufacture of which quartz is used.
5. What is water glass? How is it made, and for what purpose is it used?
6. How are clay soils formed? Why are they desirable for agricultural purposes?
7. Discuss the occurrence of silica in natural waters, also in plants and animals.
8. What is the action of hydrofluoric acid upon silicates?

CHAPTER IX

CARBON AND ITS COMPOUNDS

WHEN a porcelain plate is held in the flame of a candle, carbon deposits on the plate in the form of *soot*, which is also called **lampblack**. When in the process of baking potatoes, apples, bread, or meats they are left too long in a somewhat overheated oven, only charred masses which are largely *carbon* remain. In fact, *all animal or vegetable matter when similarly heated yields residues of carbon, showing that the latter element is an essential constituent of all living beings.* Such charred remains of animal or vegetable matter always weigh much less than the original material, which is due to the fact that a large portion of the latter has been volatilized during the process of heating. What remains is very largely carbon plus the mineral constituents, *i.e* the ash. *On continued heating in the air the carbon in the charred mass gradually burns off, and there is finally left nothing but the ash.*

FIG. 22. — Making charcoal. A vertical section through the center of a pile.

By heating wood out of contact with the air in ovens **charcoal** is produced. When coal is similarly heated, **coke** is formed; and when bones are thus treated, **bone black** results. In like manner **blood charcoal** may be made from blood. Charcoal is generally

produced by piling up wood in a suitable manner, covering the whole with earth to exclude undue access of air, and then setting fire to the wood. Thus a portion of it burns, and the rest is merely charred because it is heated practically out of contact with the air. *Charcoal, coke, and bone black always contain, besides carbon, the mineral substances,* i.e. *the ash, of the original material that was charred.*

FIG. 23. — Charcoal burning in Bavaria.

Diamond *is crystalline carbon. It is the hardest substance known.* When pure it is colorless and has a high index of refraction, for which reason it is greatly prized as a gem. *Less desirable pieces of diamond are used for making drills to drill rocks.* **Graphite** also occurs in nature and is not infrequently in a fairly pure state. It is soft, black, and "soapy" to the touch. It is used as a lubricant, also for making "lead" pencils. Graphite is also termed **plumbago,**

which name originated because it was formerly thought that the substance contained lead. Graphite is now made artificially by heating carbon very highly out of contact with the air in the electric furnace, and then allowing it to cool very slowly. Under these conditions the amorphous carbon is changed to graphite. The latter may now be obtained in pieces of almost any size desired. These can readily be worked into any form required by means of lathes, milling machines, or bench tools. Much of this artificial graphite is now used in making pencils, electrodes for electrolytic work, resistances, etc.

Charcoal will absorb gases, and hence it is frequently used to take odoriferous gases out of cisterns, pits, vaults, etc. The cause of such gases ought in all cases to be removed, however, if possible. To hang a bag full of charcoal in an ill-smelling cistern, when the latter contains decaying organic matter that ought to be taken out, is obviously not the proper way of dealing with the problem. In the sick room charcoal dressings are at times employed to absorb fetid odors issuing from ulcers. Such dressings need frequent renewal to be efficient.

Bone black consists of only about 8 to 12 per cent carbon, the remainder being largely calcium phosphate. Bone black absorbs many coloring matters from solutions, and hence is used in sugar refining to remove the brown color from the sugar solutions. Blood charcoal is more expensive than bone black. It absorbs coloring matter better, and consequently is often employed in the chemical laboratory in experimental work. *Like charcoal, bone black and blood charcoal also absorb odors.*

Coal *represents the fossil remains of plants of the carboniferous and other geological ages.* By the gradual loss of water and other volatile matter, like marsh gas, hydrogen,

etc., these vegetable remains have gradually been transformed to coal. *Wood fiber, peat, brown coal, soft coal or bituminous coal, and anthracite or hard coal represent a series of gradations in the gradual process which has taken place in nature in the production of coal.* Beginning with wood fiber, by gradual loss of oxygen, hydrogen, and some carbon as well, the other products just enumerated may be regarded as having been formed. So **anthracite** generally consists of about 95 per cent carbon, the rest being ash plus a few per cent of volatile matter. In **soft coal** there generally is about 80 per cent carbon and a very much higher amount of volatile matter than in hard coal. Soft coal is consequently used for making **coal gas** for heating and illuminating purposes. In this process the coal is heated out of contact with the air in iron retorts. In the latter finally coke remains, which also contains the mineral content or ash of the coal. The following represents the composition of a fairly typical sample of coal gas: methane or marsh gas, 34.5 per cent; hydrogen, 49 per cent; carbon monoxide, 7.2 per cent; nitrogen, 2.3 per cent; oxygen, nil; carbon dioxide, 1.1 per cent; illuminants, benzene, etc., 5 per cent. Ammonia is also always formed in the manufacture of coal gas, as already stated in Chapter IV, but it is washed out of the gas before the latter is stored and admitted to the gas mains. All of the ammonia and ammonium salts of commerce are obtained from the gas liquors.

When coal gas burns from an ordinary jet, a luminous flame results which deposits soot on a plate that is held in it. *The luminosity of the flame is caused by the incandescent particles of carbon that are present in the flame.* A hydrogen flame is colorless, but can be made luminous for a time by carefully blowing fine particles of soot into it. When sufficient oxygen is supplied to a coal gas flame, its luminosity

disappears because the carbon in the flame is all consumed; for the latter reason such a flame is hotter than the luminous one. The **Bunsen burner**, Fig. 24, is an arrangement for burning gas by first mixing it with a sufficient amount of

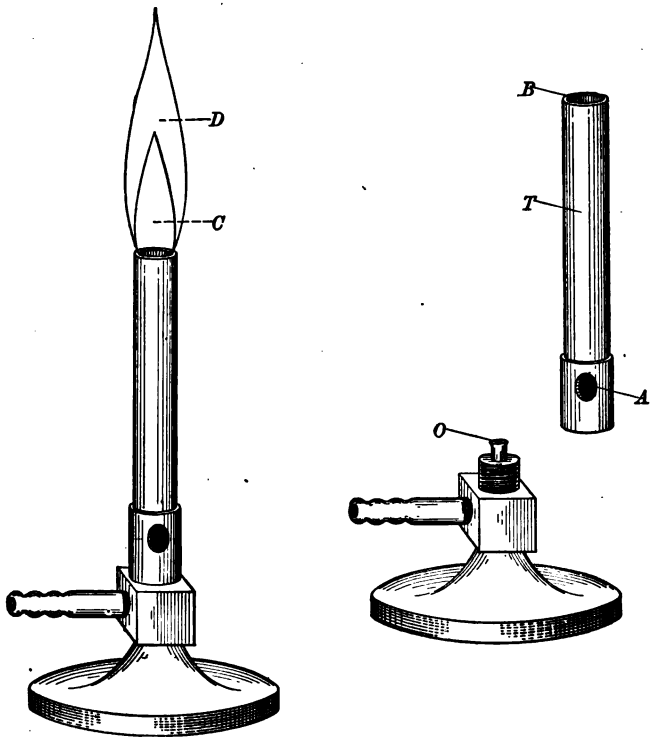


FIG. 24. — The Bunsen burner and its flame.

air so that more perfect combustion is secured. The gas flows from the small orifice *O* and air enters at the opening *A*. In the tube *T* both air and gas mix well, and this mixture is then ignited at *B*, the upper end of *T*. The gas burns with a flame that has an inner blue cone *C* and an outer zone *D* which is non-luminous. The inner zone contains

unconsumed gas, while in the outer zone practically complete combustion takes place. This is the hottest part of the flame. All gas flames intended for heating purposes are arranged on this principle. So, for example, *the burners in the laboratories, the gas stoves and ranges, the gasoline heaters, etc., are all arranged so that the gas is first well mixed with air and then the mixture is burned from a jet.* In the case of *blast lamps* air is blown into the gas by pressure from some source.

A flame results only when a gas burns. Charcoal or coke never burn with a flame, for in their production all gases were expelled from them. Charcoal or coke in burning become red hot and glow till they are consumed so that only ash remains.

When carbon burns in the air or in oxygen in excess, carbon dioxide results, thus:

$$\underset{\text{carbon}}{C} + \underset{\text{oxygen}}{O_2} = \underset{\text{carbon dioxide}}{CO_2}$$

When, however, the amount of oxygen is limited and an excess of carbon is present, **carbon monoxide** is formed, thus:

$$\underset{\text{carbon}}{C} + \underset{\text{oxygen}}{O} = \underset{\text{carbon monoxide}}{CO}$$

Carbon monoxide is a gas which burns with a blue flame, yielding carbon dioxide, thus:

$$\underset{\text{carbon monoxide}}{CO} + \underset{\text{oxygen}}{O} = \underset{\text{carbon dioxide}}{CO_2}$$

Carbon monoxide is poisonous. When breathed it unites with the hemoglobin of the blood and produces death. *Carbon monoxide is the most dangerous ingredient of coal gas.* Its flame is quite hot. In "**water gas,**" *which is a mixture of*

hydrogen and carbon monoxide produced by passing steam over white-hot coke, thus:

$$\underset{\text{carbon}}{C} + \underset{\text{water}}{H_2O} = \underset{\text{water gas}}{CO} + H_2,$$

we have then a gas which is excellent for heating, but not at all suitable for illuminating purposes. By "*enriching*" this gas with benzene, or gas produced from petroleum oils, it may be used as illuminating gas. With the **Welsbach mantle lamp** water gas may be used very well, for here a non-luminous Bunsen flame is employed, over which the mantle, consisting of a network of 1 per cent cerium oxide and 99 per cent thorium oxide, is placed. This "stocking" is heated to incandescence and emits a brilliant light. **Producer gas** consists of 28 to 30 per cent carbon monoxide, 63 per cent nitrogen, and minor amounts of carbon dioxide. It is made by passing air over red-hot coke. The gas is readily made and is frequently used as a source of heat in industrial processes. The nitrogen it contains is, of course, a drawback, for it acts simply to dilute the carbon monoxide.

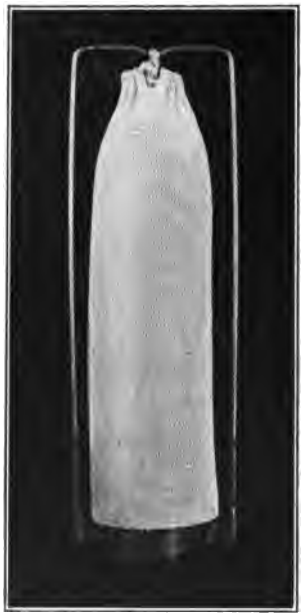


FIG. 25. — A Welsbach mantle.

Carbon dioxide, as already stated, is contained in the atmosphere, and from this source the carbon that is found in living beings is really obtained. *In the green leaf of the plant in the sunlight carbon dioxide from the air and moisture*

react with each other, forming starch and eliminating oxygen, thus:

$$\underset{\text{carbon dioxide}}{6\ CO_2} + \underset{\text{water}}{5\ H_2O} = \underset{\text{oxygen}}{6\ O_2} + \underset{\text{starch}}{C_6H_{10}O_5}$$

Starch *is an article of food. It is present in large quantities in all cereals, potatoes, and other vegetables.* As these are eaten the starch is changed to sugar in the alimentary canal, passes through the walls of the latter into the blood, and is utilized in building up tissues of the body. The latter as we live are again continually slowly oxidized by the oxygen which passes through the walls of the lungs into the blood as we breathe. *Thus carbon dioxide is formed and this is exhaled with every breath.* In this way carbon is returned to the air in the form of carbon dioxide and the **carbon cycle** has been completed. The plant can now again take up the carbon dioxide and transform it into starch. The latter can again be eaten, assimilated, oxidized to carbon dioxide in the body, and exhaled, and so on. *The energy that keeps this process going obviously comes from the sun without whose light and warmth the plant could not live and continue its work of transforming carbon dioxide and water into starch and oxygen.*

Carbonated waters are waters charged with carbon dioxide, the latter gas being placed under pressure over water in a closed tank. When such pressure is released from the water, a considerable amount of the gas escapes in bubbles. *Natural springs whose waters are charged with carbon dioxide occur in nature.* So, for example, at Colorado Springs the waters are highly carbonated. **Soda water** *is water charged with carbon dioxide.* It is so called because the gas employed was formerly made by the action of an acid on baking soda, thus:

$$\underset{\text{baking soda}}{2\ NaHCO_3} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{2\ CO_2}$$

Carbonated waters are refreshing to the taste and are consequently rightfully highly esteemed. Carbon dioxide from the fermentation industries is now commonly stored in steel cylinders and used for making carbonated beverages of all kinds. Since carbon dioxide occurs in the air and since water dissolves carbon dioxide, the latter is present in all natural waters.

Carbonates *are salts of carbonic acid.* Carbon dioxide is really *carbonic acid anhydride,* that is, carbonic acid minus water. **Carbonic acid** is H_2CO_3, but it has never been isolated, for it decomposes readily into carbon dioxide and water, thus:

H_2CO_3	=	CO_2	+	H_2O
carbonic acid		carbon dioxide		water

The salts of carbonic acid are, however, quite common and stable under ordinary conditions. They are also of great importance. So **limestone** and **marble** consist of almost pure calcium carbonate, $CaCO_3$. Magnesium limestone or **dolomite** consists of magnesium and calcium carbonate, $MgCO_3 + CaCO_3$. *These are valuable as building materials,* (1) *as building stones and* (2) *for making building lime.* In the process of making lime the limestone is heated in a kiln out of contact with the air. Thus carbon dioxide is expelled, and **lime,** *calcium oxide,* CaO, remains:

$CaCO_3$ on heating	=	CaO	+	CO_2
calcium carbonate limestone		calcium oxide lime		carbon dioxide

In *slaking the lime,* as for building purposes, making whitewash for walls, or producing liquids for spraying trees, etc. calcium hydroxide results and considerable amounts of heat are evolved as the chemical union of lime and water proceeds, thus:

CaO	+	H_2O	=	$Ca(OH)_2$
calcium oxide lime		water		calcium hydroxide or slaked lime

When this slaked lime is mixed with sand, **mortar** *is obtained, whose hardening or "setting" proceeds because water dries out*

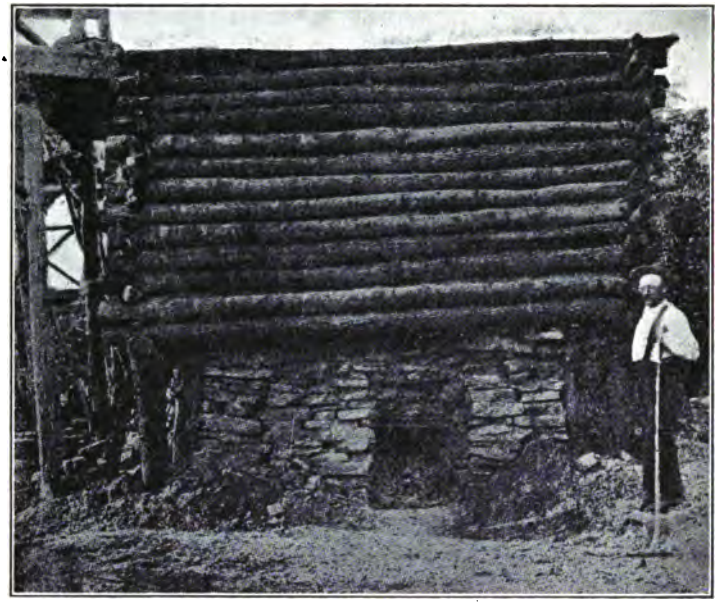


FIG. 26. — A homemade lime kiln.

and at the same time the carbon dioxide of the air again acts on the slaked lime, forming calcium carbonate again, thus:

$$\underset{\substack{\text{calcium hydroxide}\\ \text{slaked lime}}}{Ca(OH)_2} + \underset{\text{carbon dioxide}}{CO_2} = \underset{\text{calcium carbonate}}{CaCO_3} + \underset{\text{water}}{H_2O}$$

The crystals of calcium carbonate tightly hold the grains of sharp sand in place, thus forming a solid mass. It is clear that *in order to set well, lime mortar must have plenty of air so that it may secure the necessary carbon dioxide therefrom.* Moreover, the air should be fairly dry in order that the water may dry out of the mortar. *Plastering and bricklaying in cold, damp weather does not produce the best results.*

In many sands and soils grains of limestone, calcium carbonate, are found. *The shells of oysters, clams, and other shellfish consist largely of calcium carbonate.* Indeed, it is quite probable that the large beds of limestone and marble (which is merely a purer form of limestone) have been formed from the remains of mollusks, whose shells have been deposited from the sea in deep layers which have later, under heat and pressure, been transformed to the crystalline state.

Carbon is further found in nature in **petroleum**, *a liquid which consists of a mixture of compounds of carbon and hydrogen.* Compounds of the latter elements are called **hydrocarbons.** It is not easy to separate from petroleum in pure form any one of the mixture of hydrocarbons of which it consists. However, it is not at all difficult to separate petroleum into groups of hydrocarbons whose boiling points are fairly close together. This is actually done in practice by the process of **fractional distillation.** *Petroleum is distilled, and separate portions passing over between certain temperatures are collected in different receivers.* In this way there are obtained from petroleum: **rhigolene**, which boils at about room temperature; **petroleum ether**, boiling between 50° and 60° C.; **gasoline**, boiling between 70° and 90° C.; **naphtha**, boiling from about 90° to 120° C.; **benzine**, boiling from 110° to 140° C.; and **kerosene**, boiling between 150° and 300° C. Above the latter temperature heavy oil passes over as the distillation proceeds. This is used for lubricating purposes. After the **lubricating oils** have been distilled off, **vaseline** passes over, and finally **paraffine** is prepared from the residue that remains in the retort. *The oils prepared from petroleum are often called mineral oils to distinguish them from the fats and oils of plant or animal origin.* The fats and oils from the latter sources are *not* hydrocarbons. They are compounds of carbon, hydrogen, and oxygen. For the most part *they are*

salts in which glycerine is the base and oleic, palmitic, and stearic acids are the acids. By heating such fats with strong bases like caustic potash or caustic soda **glycerine** is set free, and the sodium salt of the fatty acid, a **soap**, is formed. So for example:

$$\underset{\substack{\text{glycerine oleate}\\\text{cottonseed oil}}}{(C_{17}H_{33}COO)_3C_3H_5} + \underset{\text{caustic soda}}{3\,NaOH} = \underset{\substack{\text{sodium oleate}\\\text{castile soap}}}{3\,C_{17}H_{33}COONa} + \underset{\text{glycerine}}{C_3H_5(OH)_3}$$

Mineral fats and oils are hydrocarbons and consequently cannot be transformed into soaps by boiling with lye. It is thus easy to distinguish between mineral oils and those of plant or animal origin.

Petroleum products cannot serve as foods. There is no hydrocarbon which can be used for food. The lower boiling hydrocarbon oils from petroleum are used as fuels, for heating, and for propelling internal combustion engines. When completely burned they form carbon dioxide and water. Hydrogen burns more readily than carbon, and *the richer in hydrogen a hydrocarbon is, the more volatile it is and the more readily it burns. For these reasons petroleum ether, gasoline, naphtha, and benzine are more suitable than kerosene for running so-called gasoline engines,* although kerosene is used at present with a fair degree of success, especially in those engines in which that hydrocarbon is transformed into a very fine mist by mechanical means. This mist when mixed with air and subjected to the action of the electric spark in the engine cylinder explodes fairly readily. It should be remembered that *all gases which burn in the air will also form mixtures with air which are explosive when ignited. Upon this fact the running of gas and gasoline engines depends.* The fuel is in each case volatilized and mixed with air, and the mix-

ture is then exploded in the cylinder of the engine by ignition.

Acetylene is another hydrocarbon which has come into prominent use, particularly for lighting purposes. It is a gas at ordinary temperatures and is readily prepared by the action of **calcium carbide**, a stonelike solid, upon water, thus:

$$\underset{\text{calcium carbide}}{CaC_2} + \underset{\text{water}}{2\,H_2O} = \underset{\text{acetylene}}{C_2H_2} + \underset{\text{calcium hydroxide slaked lime}}{Ca(OH)_2}$$

The calcium carbide is obtained by heating lime and coke together in the electric furnace out of contact with the air, thus:

$$\underset{\text{lime}}{CaO} + \underset{\text{coke}}{3\,C} = \underset{\text{calcium carbide}}{CaC_2} + \underset{\text{carbon monoxide gas}}{CO}$$

On account of its high carbon content, *acetylene has great illuminating power* and is consequently frequently used for lighting purposes, especially on automobiles, in running projection lanterns, etc.

From petroleum oils, too, gas of high illuminating power may be obtained by " cracking " the oils, which consists of having them come into contact with red-hot stone surfaces out of contact with the air. These **oil gases** are now produced in many municipal gas plants. They are also put on the market compressed in steel cylinders. **Pintsch gas,** which is used for lighting railway cars, is a rich oil gas of this character. **Blaugas,** named from its inventor Blau, is also a similar oil gas.

Hydrocarbons are very numerous and they may be regarded as the mother substances from which all other compounds of carbon are derived. The simplest hydrocarbon is **methane,** or **marsh gas,** CH_4. It is called marsh gas because it issues from the marshes on warm summer days as the vegetable matter which is always present on the bottom of such waters

slowly decays. The gas may also be prepared artificially, by heating sodium acetate with a caustic alkali, thus:

$$\underset{\text{sodium acetate}}{CH_3COONa} + \underset{\text{sodium hydroxide}}{NaOH} = \underset{\text{sodium carbonate}}{Na_2CO_3} + \underset{\text{methane}}{CH_4}$$

Marsh gas will burn in the air, thus:

$$\underset{\text{marsh gas}}{CH_4} + \underset{\text{oxygen}}{2\,O_2} = \underset{\text{carbon dioxide}}{CO_2} + \underset{\text{water}}{2\,H_2O}$$

Mixed with air it forms an explosive gas mixture, which could be used for running an internal combustion engine.

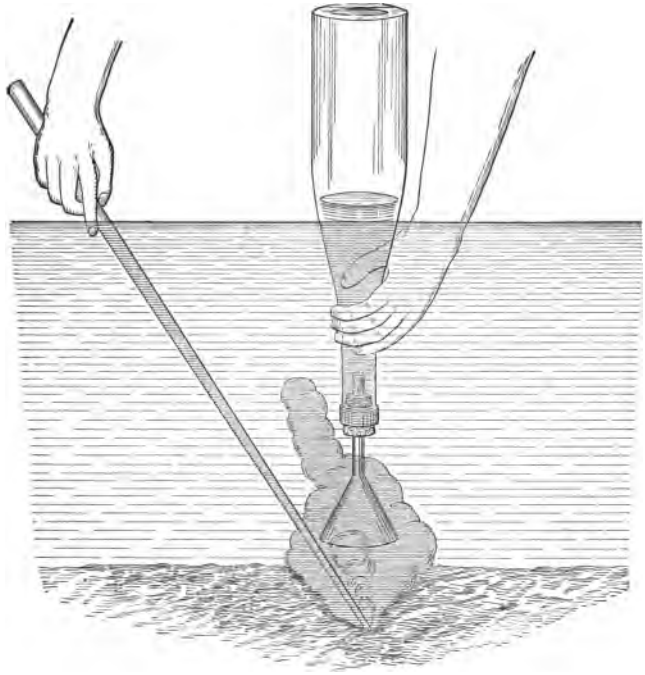


FIG. 27. — Collecting marsh gas from the bottom of a ditch.

From marsh gas a variety of products may be obtained. Thus by treatment with chlorine there may be obtained

successively, methyl chloride CH_3Cl, methylene chloride CH_2Cl_2, **chloroform** $CHCl_3$, and **carbon tetrachloride** CCl_4, viz.:

$$CH_4 + Cl_2 = HCl + CH_3Cl$$
$$CH_3Cl + Cl_2 = HCl + CH_2Cl_2$$
$$CH_2Cl_2 + Cl_2 = HCl + CHCl_3$$
$$CHCl_3 + Cl_2 = HCl + CCl_4$$

Of these compounds chloroform and carbon tetrachloride are by far the most important ones. Both are liquids at ordinary temperatures. *Chloroform boils at 61° C., is heavier than water, has a rather agreeable odor, and is used as an anæsthetic.* The analogous iodine compound, **iodoform**, CHI_3, *is a yellow crystalline solid of rather disagreeable odor.* It is used in surgery as a dressing for wounds, especially when pus has formed. *Carbon tetrachloride is also a colorless, heavy liquid which does not dissolve in water. It boils at 76° and it is not inflammable,* which, of course, is to be expected, considering that it consists of carbon and chlorine. *Carbon tetrachloride is an excellent solvent for fats, and hence is often used for removing grease spots from clothes,* for which purpose it is to be recommended, for it does not form dangerous explosive mixtures with the air like gasoline, for example, which is also used in cleaning clothes. However, carbon tetrachloride is much more expensive than gasoline. The trade name "**carboneum**" is sometimes given to carbon tetrachloride. *The latter compound is also used in fire extinguishers of various forms, particularly those of the syringe type, which are often sold at exorbitant prices.* Carbon tetrachloride extinguishes fires because it evaporates and crowds the air away from the burning substances, thus making it impossible for combustion to proceed further. During the evaporation of the carbon tetrachloride heat is absorbed, and so the temperature of the burning substances is lowered and this also tends to check the fire.

Fires *may be extinguished in the following ways:* (1) *by cutting off the supply of air from the burning materials,* (2) *by lowering the temperature of the materials on fire so that further combustion is impossible.* Now when water is thrown upon an ordinary fire the latter is extinguished for two reasons; namely, (1) *the water lowers the temperature,* and (2) *it also forms clouds of water vapor which crowd the air way.* It is obvious that when oils, which will float on water, are burning, the addition of water cannot prevent access of air to such a fire, for the oils will come on top of the water and continue to burn, the water serving rather as a means of spreading the burning oil. Many incipient *fires can be put out by covering them with blankets, rugs, earth, etc., and thus shutting off the air. Burning oils are extinguished by means of carbon dioxide and carbon tetrachloride.* Still other chemicals have been suggested for this purpose, but they have not come into common use. *Carbon dioxide is one of the very best means known for putting out small fires.* It may be delivered from cylinders containing compressed carbon dioxide, or may be generated by the action of sulphuric acid upon a saturated solution of baking soda, sodium bicarbonate, $NaHCO_3$. The *fire extinguishers* that are to be inverted and are then ready for use, Fig. 28, contain a solution of sodium bicarbonate and a glass bottle in which sulphuric acid has been placed. This bottle is loosely stoppered with a lead stopper. On inverting the entire apparatus the sulphuric acid bottle is inverted, its stopper drops out, the acid mingles with the sodium bicarbonate solution, carbon dioxide is generated, and the pressure that thus results

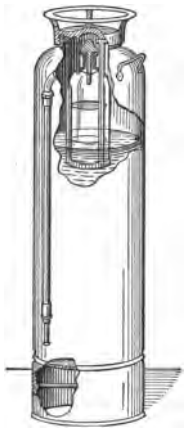


FIG. 28. — A fire extinguisher partly opened to show its construction.

is sufficient to cause a stream of the saturated effervescing solution to be delivered from the nozzle. This directed upon the base of the flames cools the burning material and at the same time the carbon dioxide liberated crowds the air away and thus extinguishes the fire. It need hardly be stated that small fires should be promptly put out, and the means for doing so should be on hand wherever there is special danger. Large conflagrations are at times well-nigh impossible to cope with because of the mass of water that would be required to prevent access of air and to lower the high

FIG. 29. — A fire extinguisher in action.

temperatures below the kindling point. In such cases the best that can be done is to prevent the spread of the fire by removal of combustible material to which there is immediate danger that the flames will spread.

When methyl chloride, CH_3Cl, is treated with caustic soda, common salt splits off and simultaneously *methyl hydroxide*, also called **methyl alcohol** or **spirits of wood**, is formed, thus:

$$\underset{\substack{\text{methyl chloride}\\ \text{H}}}{CH_3Cl} + \underset{\text{caustic soda}}{NaOH} = \underset{\text{common salt}}{NaCl} + \underset{\text{wood alcohol}}{CH_3OH}$$

Methyl alcohol is called **wood alcohol** because it is commercially produced by the dry distillation of wood, that is to say by heating wood out of contact with the air, in which process a variety of other products like **pyroligneous acetic acid, creosote**, etc., are also formed. *Methyl alcohol is poisonous. It is used for fuel and for making various chemical substances.* So, for instance, upon oxidation of methyl alcohol **formaldehyde** is obtained, thus:

$$\underset{\text{methyl alcohol}}{CH_3OH} + \underset{\text{oxygen}}{O} = \underset{\text{formaldehyde}}{HCOH} + \underset{\text{water}}{H_2O}$$

Formaldehyde is a pungent, poisonous gas which is soluble in water. Its 40 per cent solutions are sold on the market as **formaline.** It serves as an antiseptic in fumigating rooms, and its solutions serve in treating seeds to free them from injurious fungi, etc., before planting. So, for example, potatoes that are infested with **potato scab** are advantageously treated with formaldehyde solution before planting. The use of formaldehyde as a preservative in milk and other foods and drinks is forbidden by law because the substance is poisonous.

On further oxidation of formic aldehyde, **formic acid, a** fairly powerful acid, may be obtained, thus:

$$\underset{\text{formaldehyde}}{HCOH} + \underset{\text{oxygen}}{O} = \underset{\text{formic acid}}{HCOOH}$$

But one of the hydrogens of this acid is replacable by metals, and when this has been accomplished the resulting salts are the **formates.** Formic acid was formerly made by the distillation of red ants, in which it occurs.

Ordinary **alcohol,** also called **grain alcohol** or **spirits of wine,** is made commercially by the fermentation of sugar by means

of yeast. As the yeast plant, Fig. 30, lives it produces carbon dioxide from the sugar, thus:

$$\underset{\text{grape sugar}}{C_6H_{12}O_6} + \text{yeast} = \underset{\text{carbon dioxide}}{2\,CO_2} + \underset{\text{alcohol}}{2\,C_2H_5OH}$$

FIG. 30. — Yeast cells.

Like wood alcohol, grain alcohol may serve as a fuel. Both substances burn with a hot, blue flame, forming carbon dioxide and water. On careful oxidation ordinary alcohol forms first **acetic aldehyde** and then **acetic acid**, thus:

$$\underset{\text{ordinary alcohol}}{CH_3.CH_2.OH} + O = \underset{\text{acetic aldehyde}}{CH_3COH} + \underset{\text{water}}{H_2O}$$

$$\underset{\text{acetic aldehyde}}{CH_3COH} + \underset{\text{oxygen}}{O} = \underset{\text{acetic acid}}{CH_3COOH}$$

The oxidation of alcohol to acetic acid is accomplished commercially by means of an organism called mother of vinegar. The dilute alcohol is allowed to trickle over beech wood shavings contained in a vat, and in the presence of air and the acetic acid organism the alcohol is oxidized to **vinegar.** *The latter is a solution of about 4 per cent acetic acid.* From

this by the action of bases a series of salts called **acetates** is formed. Of the metallic acetates **lead acetate**, $Pb(CH_3CO_2)_2$, is perhaps the most common in practice, though **sodium acetate**, $Na \cdot CH_3CO_2$, and **copper acetate**, $Cu(CH_3CO_2)_2$, are also

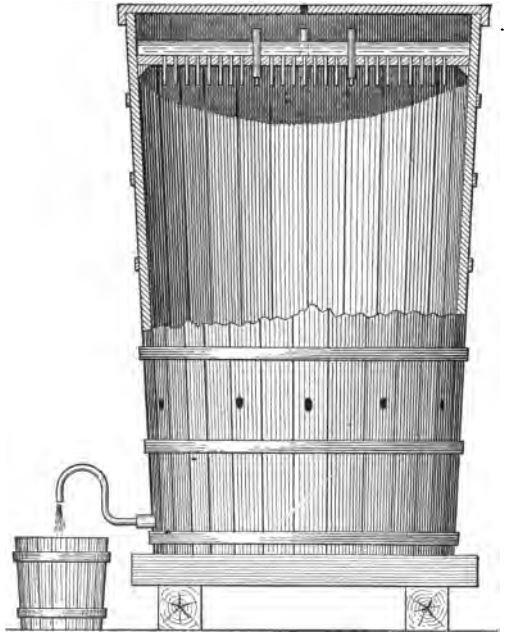


FIG. 31. — Making vinegar by the quick vinegar process.

often used. The latter, it will be recalled, forms Paris green when united with copper arsenite. *Vinegar is also formed by the fermentation of sugar and subsequent oxidation of the resulting alcohol in fruit juices, like cider, for example.* By adding to cider an antiseptic like boric acid, benzoic acid, or sodium benzoate, fermentation may be prevented, but as these substances are deleterious to health, the practice is not to be recommended. *Cider or grape juice may be kept unfermented*

if it is carefully heated to destroy the microörganisms it contains, and then, while still hot, sealed in air-tight bottles.

Denatured alcohol is grain alcohol which has been rendered unfit for drinking purposes by adding about 10 per cent of wood alcohol, pyridine, C_5H_5N, or other *poisonous* liquid. *Denatured alcohol is sold duty-free and is used as a fuel and for manufacturing purposes.*

Whiskeys and **rum** contain from 45 to 65 per cent of alcohol, **wines** from 8 to 20 per cent, and **beers** from 3 to 5 per cent. *As beverages, alcoholic liquids are now recognized as neither necessary nor desirable for the best of health.*

There are many other alcohols besides wood alcohol and grain alcohol; however but few of them are in common use. **Glycerine** is really an alcohol. Its composition is expressed by the formula $C_3H_5(OH)_3$. *All alcohols contain the* OH *group. They may all act as bases, reacting with acids to form salts and water.* So, for instance,

$$\underset{\text{alcohol}}{C_2H_5OH} + \underset{\text{acetic acid}}{CH_3COOH} = \underset{\text{ethyl acetate}}{CH_3COOC_2H_5} + \underset{\text{water}}{H_2O}$$

It will be seen at once that this reaction is similar to that when sodium hydroxide acts on acetic acid:

$$\underset{\text{sodium hydroxide}}{NaOH} + \underset{\text{acetic acid}}{CH_3COOH} = \underset{\text{sodium acetate}}{CH_3COONa} + \underset{\text{water}}{H_2O}$$

Salts, like ethyl acetate, in which there is a hydrocarbon radical which plays the rôle of a metal in that it has basic properties, are called **ethereal salts** or **esters.** *Esters are quite common in nature.* So, for example, **wintergreen oil, methyl salicylate,** is a salt of **salicylic acid,** in which methyl, CH_3, the radical from wood alcohol, acts as a base. **Banana oil, amyl acetate,** $CH_3COO \cdot C_5H_{11}$, is formed when acetic acid

and **amyl alcohol** react with each other, thus:

$$\underset{\text{acetic acid}}{CH_3COOH} + \underset{\text{amyl alcohol}}{C_5H_{11}OH} = \underset{\text{amyl acetate}}{CH_3COOC_5H_{11}} + \underset{\text{water}}{H_2O}$$

The fats are salts in which stearic, palmitic, and oleic acids act as acids and glycerine acts as the base. So in **mutton tallow** we have principally **glycerine stearate,** in **palm oil** chiefly **glycerine palmitate,** while **hog's lard, cotton seed** and **olive oils** are rich in **glycerine oleate.** *In general the fats are mixtures of these three salts of glycerine.* The softer the fat, the richer it is in oleine or glycerine oleate; the harder the fats, the richer they are in stearine or glycerine stearate. On heating the fats with caustic soda or caustic potash the corresponding salts of sodium or potassium are formed and glycerine is simultaneously produced. These sodium salts of stearic, palmitic, and oleic acids are the **hard soaps.** The corresponding potassium salts are the **soft soaps.** *Glycerine is a product of the process of saponification of fats,* for example:

$$\underset{\substack{\text{glycerine}\\\text{stearate}}}{(C_{17}H_{35}COO)_3C_3H_5} + \underset{\substack{\text{caustic soda}\\\text{lye}}}{3\,NaOH} = \underset{\substack{\text{sodium stearate}\\\text{hard soap}}}{3\,C_{17}H_{35}COONa} + \underset{\text{glycerine}}{C_3H_5(OH)_3}$$

When used in hard water, soaps form an insoluble curdy mass. This is really the **calcium soap. Hard water** contains calcium salts which react with soap, forming the insoluble calcium soap and a soluble sodium salt, thus:

$$\underset{\text{sodium stearate}}{2\,C_{17}H_{35}COONa} + \underset{\text{calcium sulphate}}{CaSO_4} = \underset{\text{calcium stearate}}{(C_{17}H_{35}COO)_2Ca} + \underset{\text{sodium sulphate}}{Na_2SO_4}$$

When hard water is first treated with washing soda, Na_2CO_3, it *may be "cleansed";* that is to say, the calcium may be precipitated in the form of an insoluble compound, calcium carbonate, $CaCO_3$, and then the water will not form a calcium soap when treated with ordinary soap. So, for example:

$$\underset{\text{calcium sulphate}}{CaSO_4} + \underset{\text{sodium carbonate}}{Na_2CO_3} = \underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{calcium carbonate}}{CaCO_3}$$

Instead of washing soda, borax may also be employed to soften the water, in which case calcium borate is precipitated. Ammonium carbonate may also serve for the same purpose. *It is best, however, to use rain water or distilled water for washing purposes and so avoid the use of soda, etc., for the process of cleansing water is always expensive, and the salts that remain in the water are hard on the clothes and also on the hands of the laundry workers.*

Ether is made by very carefully conducting alcohol into concentrated sulphuric acid heated to about 140° C. Under these conditions alcohol loses water, which is taken up by the sulphuric acid, thus:

$$\underset{\text{alcohol}}{2\,C_2H_5OH} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\substack{\text{sulphuric acid}\\ \text{hydrate}}}{H_2SO_4 . H_2O} + \underset{\text{ether}}{(C_2H_5)_2O}$$

Ether is the oxide of ethyl, whereas alcohol, it will be recalled, is the hydroxide of ethyl. Ether boils at 35° C. and has a specific gravity of 0.736. *It is very inflammable.* In this respect it is even more dangerous than gasoline. *Ether is used in medicine as an anæsthetic. It is also employed as a solvent for fats and oils,* especially in analytical chemistry. **Sulphur ethers** contain sulphur where ordinary ethers contain oxygen. Sulphur ethers have extremely nauseating odors.

Carbolic acid or **phenol**, C_6H_5OH, is really not an organic acid at all. It is an hydroxide derived from **benzene**, C_6H_6, *a hydrocarbon present in the light oil obtained by distilling coal tar* produced in the gas works. In reality, then, carbolic acid is closely related to alcohols, but since its hydroxyl hydrogen is easily replaced by metals forming so-called **phenolates**, phenol is commonly called carbolic acid. It is a solid whose crystals melt at 42° C. It turns pink on exposure to air. It is *poisonous and corrosive to the skin.* Fifteen parts of cold water dissolve about 1 part of phenol, and *the solu-*

tion serves as an antiseptic. Phenol has a very characteristic odor. **Creosote** is a mixture of **guajacol**, $C_6H_4(OCH_3)OH$, and **creosol**, $C_6H_3(CH_3)(OCH_3)OH$, produced when wood is heated with but limited access of air, as in **smoking fish, hams, bacon**, etc. These phenols penetrate into the meats and so preserve them from the ravages of microörganisms, thus preventing putrefaction. **Hydrochinone**, $C_6H_4(OH)_2$, and **pyrogallol**, $C_6H_3(OH)_3$, also called **pyrogallic acid**, are also phenols. They are used as **developers in photography.**

From the hydrocarbon benzene, C_6H_6, are derived many other useful substances, such as **aniline** $C_6H_5NH_2$, **nitrobenzene** or **oil of mirbane**, $C_6H_5NO_2$, **benzoic acid** C_6H_5COOH, whose sodium salt, **sodium benzoate** C_6H_5COONa, has been used as a preservative for foods. Practically all the **dyestuffs** that are now used for dyeing fabrics of all kinds are derivatives of benzene, C_6H_6. They are consequently often spoken of as **coal tar dyes, aniline dyes**, or coal tar coloring matters. They are excellent, and have practically completely replaced vegetable coloring matters in the arts.

Among the many other **organic acids** that exist there are the following, which are frequently met in daily life: (1) **oxalic acid** $(COOH)_2$, which is made from sawdust by oxidizing the latter with the aid of nitric acid. *It serves for removing ink and rust spots from floors and fabrics*, and is frequently used as a reducing agent in the laboratory. (2) **Lactic acid,** which occurs in sour milk. It has the formula $C_2H_4(OH).COOH$. With bases it forms **lactates.** *It is used with baking soda in making biscuits, pancakes, etc.* Its silver salt is used as an antiseptic in medicine. (3) **Malic acid,** $CH_2COOH.CH.OH.COOH$ occurs in sour apples, in mountain ash berries, and in many other fruits. (4) **Tartaric acid** $(CH.OH.COOH)_2$ occurs in grapes. Its acid potassium salt $(CH.OH.COOH)(CH.OH.COOK)$ is **cream**

of tartar and is used in baking powders. (5) **Citric acid** $(CH_2COOH)_2 . CH . OHCOOH$ occurs in lemons and other citrus fruits. (6) **Butyric acid**, C_3H_7COOH, is contained in rancid butter and gives to the latter its disagreeable odor. (7) **Valeric acid**, C_4H_9COOH, is contained in the catnip plant, and it is the odor of this acid that is so much esteemed by cats. (8) **Hippuric acid**, $C_6H_5 . CO . NH . CH_2 . COOH$ occurs in the urine of herbivorous animals.

The so-called **carbohydrates** *form a large and extremely important group of compounds of carbon.* They all contain only the elements carbon, hydrogen, and oxygen, and moreover the hydrogen and oxygen are always present in the same proportions by weight as in water, whence the name carbohydrate. *The carbohydrates consist of* (1) *the celluloses,* (2) *the starches,* (3) *the gums and dextrines,* (4) *the sugars.*

Wood, cotton, straw, hemp, linen, etc., when burnt leave behind a certain amount of ash, but aside from this, which represents their mineral content, they consist of almost pure cellulose. **Cellulose**, then, is one of the most widely distributed compounds in nature, being the material out of which the cell walls of all plants are made. The composition of cellulose is expressed by the formula $(C_6H_{10}O_5)_x$. Cellulose is insoluble in water, and when completely burned forms carbon dioxide and water. In the form of hay, cornstalks, straw, etc., *cellulose serves as a food for cattle, horses, and other herbivores,* but human beings and carnivorous animals can *not* use this material for food. It is evident, however, that *for fuel, clothing, and shelter, cellulose in various forms is invaluable to mankind.* **Paper** is made from rags, straw, wood pulp, etc.; it therefore consists of fibers of cellulose that are matted together. **Filter paper** and **blotting paper** are unsized papers; whereas **writing paper** and many other papers that have a hard, smooth, finished surface contain rosin which has been

melted and rolled into the sheets by means of hot rollers. The surface thus finished does not absorb ink so readily, and hence serves well for writing purposes.

Nitrates of cellulose may be formed by treating cellulose with a mixture of nitric and sulphuric acids. These **nitrocelluloses** are true esters of cellulose, for on saponifying them with caustic soda, sodium nitrate and cellulose are formed. **Guncotton** is cellulose hexanitrate, $C_{12}H_{14}(NO_3)_6O_4$. It looks much like ordinary cotton, but it is harsher to the touch and on ignition it burns very rapidly and quietly, forming no smoke. *Guncotton can be made to explode violently by means of a cap of fulminating mercury.* The latter is a dangerous white powder that results when mercury is acted upon by nitric acid in presence of alcohol. *Guncotton, then, serves as an explosive.* It may be used alone or together with **nitroglycerine**. The latter is a viscous yellowish liquid. It is the nitrate of glycerine and is produced when glycerine is acted upon by a mixture of concentrated nitric and sulphuric acids. When absorbed in diatomic earth and molded into sticks, it forms **dynamite**. Nitroglycerine, like guncotton, may be exploded by means of fulminating mercury, but it may also explode from other shocks and it is consequently a very dangerous compound. With the aid of vaseline and a solvent like acetone, $(CH_3)_2 . CO$, guncotton and nitroglycerine are worked into threads which are used as **smokeless gunpowder**. *Guncotton, nitroglycerine, and dynamite are used in the arts for blasting purposes.* They have also supplanted black gunpowder in military operations. While gun cotton is not soluble in a mixture of alcohol and ether, cellulose tetranitrate and pentanitrate are soluble in this solvent and form a viscous solution that is known as **collodion.** *It is used for making films in photography, and in medicine it is employed in dressing wounds.* On evaporation of the

alcohol and ether the nitrocellulose remains as a tough transparent film. **Celluloid** *is made of guncotton and camphor.* It is therefore a *highly combustible material,* and this fact should always be borne in mind in using it, so that serious accidents may not result on subjecting it to heat and especially to flames.

Starch, too, consists of carbon, hydrogen, and oxygen, and these are present in the same proportion by weight as in cellulose; but they are combined in a different way than in the latter. Starch, then, has the formula $(C_6H_{10}O_5)_x$. It will not dissolve in water, but when a little of it is rubbed fine in about twice its volume of cold water and then ten to twenty times that volume of boiling water is added to the mixture with constant stirring, **starch paste** results which is used for starching clothes; it also frequently serves as an adhesive. When starched clothes are afterwards ironed, a gloss is produced because under the hot iron the starch is converted into dextrine which gives the luster to the linen fibers. *Starch occurs in all plants, especially in potatoes and other tubers and fleshy roots, but also in grains like rice, wheat, rye, oats, barley, corn, etc.* The supply of starch which all seeds contain serves to nourish the young plant till its roots and tops have developed sufficiently to draw sustenance from the soil and the air. In this process starch is gradually changed to sugar, and the latter, being soluble in water, is utilized by the growing embryonic plant. *Starch is a most important article of food for man and animals as well.* **Flour** consists of about 70 per cent starch, together with 10 per cent **gluten** (which is a nitrogen-bearing substance that is akin to the albumins found in the white of egg and hence is also valuable as food) and small quantities of mineral matter (*i.e.* ash), water, and sugar. When heated to 210° C. starch is converted to **dextrine,** $C_6H_{10}O_5$, which is a colorless amorphous substance. On

treatment with water it yields a sticky mass, and hence is very commonly used as a cheap adhesive gum. *On heating starch with dilute sulphuric acid, it is converted into glucose.* The sulphuric acid may afterwards be removed by treatment with lime, for thus insoluble calcium sulphate is produced.

Sugars *may be divided into two classes, the monoses and bioses.* The **monoses** have a composition corresponding to the formula $C_6H_{12}O_6$, and the common representatives of this group are (1) **glucose** or **dextrose**, also known as **grape sugar**, (2) **levulose**, or **fruit sugar**, also called **fructose**. The composition of the **bioses** corresponds to the formula $C_{12}H_{22}O_{11}$, and the chief common representatives of this class are (1) **cane sugar**, also called **saccharose** or **sucrose**, (2) **maltose**, or **malt sugar**, and (3) **lactose**, or **milk sugar**.

Glucose *is found in the juice of many fruits*, especially of grapes, whence the name grape sugar. Its solutions turn the plane of polarized light to the *right*, hence this sugar is also called *dextrose. When yeast is added to dilute solutions of dextrose, fermentation takes place and alcohol and carbon dioxide are formed*, thus:

$$\underset{\text{glucose}}{C_6H_{12}O_6} \quad (+ \text{ yeast}) = 2 \underset{\text{alcohol}}{C_2H_5OH} + 2 \underset{\text{carbon dioxide}}{CO_2}$$

Solutions of glucose precipitate red cuprous oxide, Cu_2O, from hot **Fehling's solution**. The latter is a strongly alkaline solution of copper sulphate and Rochelle salt, sodium potassium tartrate. *This solution is much used in testing for reducing sugars.* For example, by this means in the case of diabetic patients sugar may be detected in the urine. By heating starch with dilute sulphuric acid, glucose is prepared on a large scale, as already mentioned in connection with the consideration of starch. *Thousands of tons of glucose are thus manufactured from corn starch annually in the United*

States. Glucose is a good, wholesome food. It is mainly used in candies, table sirups, etc. It is about three-fifths as sweet as cane sugar.

Levulose or **fructose** is also found in the juice of fruits. It also occurs in honey. Its solutions turn the plane of polarized light to the *left,* whence the name *levulose;* they also reduce Fehling's solution yielding a precipitate of cuprous oxide. Furthermore this sugar, too, may be fermented with yeast, yielding alcohol and carbon dioxide.

By treating cane sugar solutions with dilute acids both dextrose and levulose are produced, thus:

$$\underset{\text{cane sugar}}{C_{12}H_{22}O_{11}} + \underset{\text{water}}{H_2O} = \underset{\text{dextrose}}{C_6H_{12}O_6} + \underset{\text{levulose}}{C_6H_{12}O_6}$$

This process proceeds quite rapidly on heating the solution gently to boiling. Since the resulting liquid turns the plane of polarized light slightly to the left, whereas the original cane sugar solution turned the plane of polarized light to the right, the process of thus changing cane sugar to dextrose and levulose is commonly termed the **inversion of cane sugar.** The latter compound does not alter Fehling's solution, whereas it will be recalled that both dextrose and levulose effect its reduction.

Cane sugar is also known by the ndmes sucrose and saccharose. **Sugar cane** contains from 15 to 20 per cent of it, whereas **sugar beets** commonly yield from 10 to 20 per cent. *In sorghum, in maple sap, in nuts, and in the blossoms of many plants saccharose abounds.* Indeed, it is very widely distributed in the vegetable kingdom. **Rock candy** is almost pure saccharose and exhibits the beautiful monoclinic prisms in which this substance crystallizes. In water it is very soluble. *One part of water will dissolve three times its weight of sugar at room temperature.* On carefully heating cane

sugar it may be melted to a colorless liquid at 160° C.; on cooling this yields so-called *barley sugar*, an amorphous glassy substance, which after a time becomes crystalline. **Caramel** is a brown substance which is obtained by heating sucrose

FIG. 32. — A field of sugar beets.

to about 200° C. Water is given off as the caramel forms. The latter is much esteemed in candies. *Yeast does not cause fermentation of cane sugar solutions at all readily.* After a time, however, fermentation with production of alcohol and carbon dioxide does occur, for yeast contains a so-called *enzyme*, **invertase**, which slowly inverts sucrose to glucose and levulose, and these then ferment. In **manufacturing sugar**, the juice obtained from sugar cane or sugar beets is treated with about a 1 per cent solution of slaked lime. This converts all acids present to calcium salts, serves to coagulate

proteins present, and at the same time, being an antiseptic, it guards against fermentation. The excess of lime is then removed by means of carbon dioxide, and the solution is decolorized by filtering through bone black, after which *it is*

FIG. 33. — A field of sugar cane.

evaporated in so-called **vacuum pans** that are heated by means of steam. On cooling, crystals of sugar separate out. These are whirled in a **centrifuge** to remove the adhering brown liquid in which they grew. The latter is sold as **molasses.** **Bagasse** is the name given to the residue of the beets or cane after the removal of their juice. *The bagasse is used as fuel, made into paper, or fed to cows. The annual production of*

sugar from beets and cane is about ten millions of tons, which shows that it is used very extensively as a food.

Maltose *or malt sugar occurs in malt.* Its solutions turn the plane of polarized light to the right, reduce Fehling's

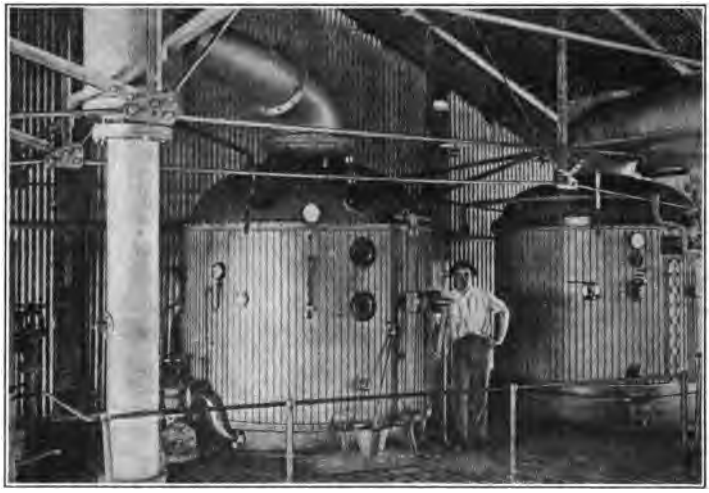


FIG. 34. — Vacuum pans used in a sugar factory.

solution, and readily ferment with yeast yielding alcohol and carbon dioxide. *Maltose is an intermediate product in the formation of alcohol from starch,* for when **diastase,** an enzyme contained in malt, acts upon starch, maltose is produced. This then may be fermented to alcohol and carbon dioxide. *Upon these facts the commercial production of fermented liquors like beer depends.* The dilute alcoholic solutions obtained by fermentation may be concentrated by fractional distillation, thus yielding whisky, gin, rum, and even alcoholic solutions of 80 per cent and still higher. Alcohol that is entirely free from water cannot be obtained by fractional distillation. When such so-called **absolute alcohol** is re-

quired, quicklime is added to alcohol that has been concentrated by distillation as far as it is profitable to do so by this means. The lime unites with the water present, but not with the alcohol, and the latter can then be distilled off.

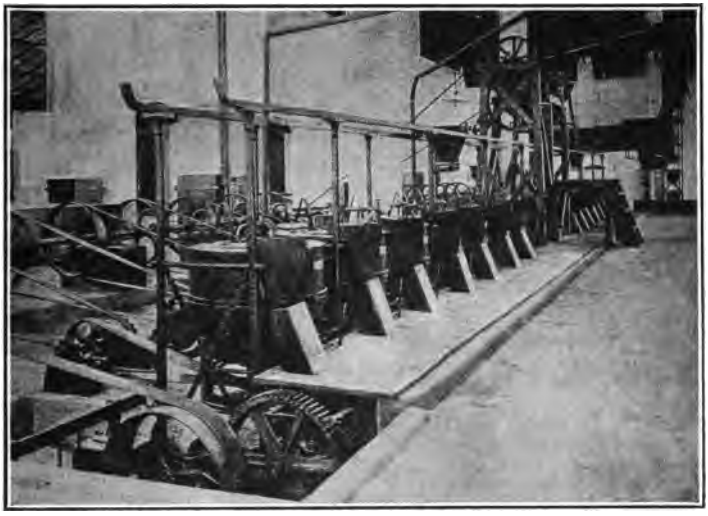


FIG. 35. — A battery of centrifuges used in drying sugar in a sugar mill.

When maltose is treated with dilute acids, dextrose only is formed.

Lactose *or milk sugar occurs in the milk of mammals.* Cow's milk contains about 5 per cent of lactose. It is not as sweet as cane sugar and far less soluble in water than the latter. Six parts of water dissolve about 1 part of lactose at room temperature. Lactose turns the plane of polarized light to the right and reduces Fehling's solution, but not as readily as maltose. Toward yeast it acts much like cane sugar, that is to say, fermentation proceeds but very slowly indeed. In this case *the products formed, however, are alcohol and lactic acid.* Yeast that is quite pure does not ferment lactose at all.

The **proteins**, *formerly called* **proteids**, *consist of carbon, hydrogen, oxygen, nitrogen, and sulphur.* They are an exceedingly important class of substances, for *when the fats, mineral matter, and water are taken from the body of an animal what is left consists entirely of proteins. No animal can live without proteins as food.* The composition of the proteins is about as follows on the average: carbon, 50.5 to 54.5 per cent; oxygen, 21 to 24 per cent; nitrogen, 15 to 17.7 per cent; hydrogen, 6.5 to 7.3 per cent; sulphur, 0.3 to 2.3 per cent; phosphorus, 0.4 to 0.8 per cent. The **nucleoproteins** so called often contain 5 to 6 per cent of phosphorus. The chemistry of the protein bodies is quite complicated, and only of recent years has notable progress toward a better comprehension of the nature of these substances been made. **Albumins** are proteins, they occur, for instance, in eggs, in muscles, in milk, in blood serum, in the seeds of plants, especially in beans, peas, and other legumes. The albumins may be coagulated by means of heat, also by means of acids. The name **albuminoids** is given to a class of protein bodies that are closely related to the albumins. So, for example, *gelatin, keratin, and elastin are albuminoids.* **Keratin** *is the main constituent of hair, hoofs, horns, feathers, nails, cuticle, etc.*, while **elastin** is found in the connective tissues. **Peptones** are formed when albumins are acted upon by **pepsin**, an enzyme, or so-called *unorganized ferment. In the change from albumin to peptone, water is added to the compounds by the action of pepsin.* This process takes place in the stomach and is important in the digestion of nitrogenous foods.

When proteins are left to putrefy, poisonous substances called **ptomaines** are produced, among which *putrescine* $C_4H_8(NH_2)_2$ and *cadaverine* $C_5H_{10}(NH_2)_2$ may be mentioned. *Poisoning caused by eating partially decayed meat, fish, or other animal food is generally due to ptomaines.*

In plants certain non-nutritive nitrogenous constituents occur, which are of importance in common life and ought consequently to be mentioned here. These substances consist of carbon, hydrogen, oxygen, and nitrogen. They have the common property that they are basic in character, that is to say, with acids they form salts. Hence these substances are called **alkaloids**. Only a few of the most important will be mentioned. Unless otherwise stated they form crystalline solids.

Quinine $C_{20}H_{24}O_2N_2$ and **cinchonine** $C_{19}H_{22}ON_2$ occur in **Peruvian bark**. *The sulphate of quinine is commonly used in treatment of malaria.* **Morphine** $C_{17}H_{19}O_3N$, **codeine** $C_{18}H_{21}NO_3$, and **narcotine** $C_{22}H_{23}O_7N$ occur in **opium**, the dried sap of the partially ripe pods of the opium poppy. *Morphine is used to produce sleep and is commonly prescribed as the hydrochloride.* **Strychnine** $C_{21}H_{22}O_2N_2$ and **brucine** $C_{23}H_{26}N_2O_4$ are found in **nux vomica**. *These are extremely bitter and very poisonous, producing death with concomitant convulsions and final muscular rigor.* **Nicotine** $C_{10}H_{14}N_2$ is found in tobacco. *It is a poisonous liquid* which boils at 247° C. **Cocaine** $C_{17}H_{21}O_4N$ occurs in coca leaves and is used as a *local anæsthetic.* **Atropine** $C_{17}H_{23}O_3N$ is found in nightshade, the deadly **Atropa belladonna.** *Oculists use this alkaloid to produce expansion of the pupil of the eye.*

FIG. 36. — An oriental poppy. From this plant opium is obtained.

QUESTIONS

1. Make a list of the various forms in which carbon occurs. How do we know that all of these are the same element carbon?

2. Mention the uses of charcoal, also of bone black.

3. What are the different varieties of coal? What is the essential difference between them?

4. How is coal gas made? What does it contain?

5. Explain how it is that ammonia is a product of the gas works.

6. How is a flame produced? Give an example.

7. What causes some flames to be luminous and others to be non-luminous? Give examples of luminous and non-luminous flames.

8. Draw a diagram of a Bunsen burner and explain it.

9. What are the properties of carbon dioxide? How is it formed? How does it occur in nature? What is it used for?

10. How is carbon monoxide formed? What are its chief characteristics?

11. What is water gas? How is it made?

12. How is starch formed? How many pounds of carbon dioxide would have to be decomposed to form 5 pounds of starch?

13. What is the so-called carbon cycle?

14. What is soda water?

15. What is lime, and how is it produced?

16. Explain the process of slaking lime.

17. What happens when lime mortar sets?

18. What is petroleum? Mention the various important petroleum products that are on the market.

19. How can a mineral oil be distinguished from one that is of plant or animal origin?

20. What is acetylene? How many grams of calcium carbide are necessary to make 100 liters of acetylene at 0° and 760 mm. pressure?

21. What is the difference between a hydrocarbon and a carbohydrate? Name five examples of each.

22. What is chloroform? For what purpose is it used?

23. What is carbon tetrachloride? What are its uses?

24. How may fires be extinguished?

25. Why should water not be used to extinguish burning oils?

26. What is wood alcohol? What are its uses?

27. How much wood alcohol would be required to produce two pounds of pure formaldehyde?

28. What is formaline, and what is it used for?

29. How is ordinary alcohol produced? What are its uses?

30. How much grape sugar would be required to produce 100 pounds of pure alcohol?

31. What is vinegar? How is it produced?

32. What is denatured alcohol? Why is it made?

33. What is glycerine? How is it produced?

34. How are alcohol and ether related chemically?

35. To what class of substances do the plant and animal fats and oils belong?

36. What is hard soap? Soft soap?

37. What is meant by the term "hard water"? How is water softened?

38. What is carbolic acid? What is it used for?

39. Why does it preserve hams and fish to smoke them?

40. What is benzene? Benzoic acid? Aniline?

41. Mention six organic acids and tell where they occur.

42. What substances are called carbohydrates and why? Where do these occur in nature?

43. Which of the carbohydrates serve as human food? Classify the carbohydrates.

44. To what animals may cellulose serve as food? Will cellulose in any form do for this purpose? Explain by a few examples.

45. What is guncotton? Dynamite? Smokeless powder? Celluloid?

46. What are the properties of starch? How test for the presence of starch?

47. What is flour?

48. Classify the sugars, and give the chief characteristics of each of the principal sugars.

49. What is a protein? Give several examples.

50. To what class of bodies do morphine, quinine, and strychnine belong?

CHAPTER X

THE METALS OF THE ALKALIES AND THE ALKALINE EARTHS.

THE metals of the alkalies are potassium, sodium, lithium, cæsium, and rubidium. The last two are of such rare occurrence that they will not be considered here. Even **lithium** is found only in small amounts. It occurs in certain micas and sometimes in a few plants like tobacco and beets. In general its properties, and also those of cæsium and rubidium, are similar to the properties of potassium and sodium. *The metals of the alkalies never occur in nature in the free or uncombined state.* They are only found with other elements in the form of salts.

FIG. 37. — Buckwheat. On the right all the elements are present; on the left all except potassium.

Plant ashes always contain potassium carbonate, K_2CO_3, also called **potash.** It has long been the practice to leach out wood ashes with water, and use the potash solution thus obtained for the purpose of making soap. **Potassium** is an element that is essential to plant and animal life. *Without potassium salts plants cannot grow.* All soils

contain potassium in the form of soluble salts, although in most cases in insufficient quantity. For this reason, it has long been the practice to return potassium salts to the land. Certain quantities get back through the application of stable manure. *Wood ashes are commonly applied to fields when potash is lacking.* **Feldspar**, which is a silicate of potassium and

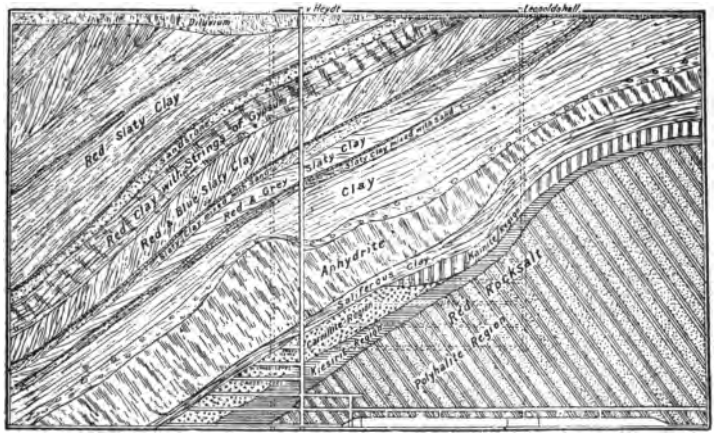


FIG. 38. — Geological formation of the Stassfurt salt beds.

aluminium, occurs in all granitic rocks and in many clay soils, and plants secure notable amounts of potash from this source, for the soil waters gradually extract potash from the feldspar, though this process is quite slow. *The largest bed of potassium salts found in nature is that at Stassfurt in Germany.* Here potassium occurs chiefly as **carnallite** $KCl \cdot MgCl_2 \cdot 6\ H_2O$ and as **kainite** $KCl \cdot MgSO_4 \cdot 3\ H_2O$. It is also found there as **sylvite, muriate of potassium**, KCl, to some extent. The potassium salt deposits at Stassfurt form layers varying in thickness from sixty to one hundred feet. They rest upon thick beds of common salt. Practically the entire supply of potassium salts of the world is derived from this enormous

deposit. However, *in small amounts potassium salts are found very widely distributed.* So potassium occurs not only in all plants, but in bones, muscles, blood, milk, albumin, and the various secretions of animals. The earth's crust contains about 2.45 per cent potassium, and in oceanic water there is about 0.04 per cent.

Caustic potash, potassium hydroxide, KOH, may be obtained by treating potassium carbonate with slaked lime, thus:

$$\underset{\text{potassium carbonate}}{K_2CO_3} + \underset{\text{calcium hydroxide}}{Ca(OH)_2} = \underset{\text{calcium carbonate (insoluble)}}{CaCO_3} + \underset{\text{caustic potash or potassium hydroxide}}{2\ KOH}$$

Large quantities of potassium hydroxide are now prepared by electrolysis of potassium chloride, muriate of potash. Thus this salt is decomposed into chlorine and potassium, and when the latter acts on water, hydrogen and potassium hydroxide are formed; the reactions involved are therefore:

$$2\ KCl \text{ by electrolysis} = 2\ K + Cl_2$$

and

$$2\ K + 2\ H_2O = 2\ KOH + H_2$$

By boiling down caustic potash solutions (which must be done in silver or iron vessels, for glass or porcelain would be strongly attacked by the lye) solid potassium hydroxide may be obtained. It is a white, hard, brittle solid which dissolves very readily in water with evolution of considerable heat. Its solutions corrode and disintegrate animal and vegetable tissues, hence the name caustic potash. *It is used in making soft soap, being placed on the market as potash lye.* On neutralization with acids, it forms various **potassium salts.** A few of the most important ones will now be described.

Potassium carbonate has already been mentioned. While it was formerly derived almost entirely from wood ashes, it is now made from the Stassfurt salts by a method similar to

that of preparing sodium carbonate by the so-called LeBlanc process (which see). Potassium carbonate is also known commercially as **pearlash.** *It is employed in the manufacture of soft soap, hard glass, and various salts of potassium.*

Saltpeter, potassium nitrate, KNO_3, is present in small amounts in all soils. It is a white crystalline salt. At ordinary temperatures 100 parts of water dissolve 13 parts of potassium nitrate, but the same amount of boiling water dissolves 247 parts of this salt. Saltpeter is prepared commercially by adding potassium chloride (from the Stassfurt deposits) to hot saturated solutions of **Chili saltpeter,** $NaNO_3$, thus:

$$\underset{\text{sodium nitrate}}{NaNO_3} + \underset{\text{potassium chloride}}{KCl} = \underset{\text{sodium chloride}}{NaCl} + \underset{\text{potassium nitrate}}{KNO_3}$$

The common salt, NaCl, is not nearly as soluble as the potassium nitrate, consequently the latter remains dissolved while the former is precipitated. From the clear *solution,* crystals of potassium nitrate are then obtained on cooling. Saltpeter is used in the manufacture of **black gunpowder,** which consists of 75 parts saltpeter, 11 parts sulphur, and 17.5 parts charcoal. When it is exploded a portion of the latter ingredient always remains unoxidized and so forms black smoke. *The use of black gunpowder in war is now a thing of the past.* Saltpeter is also used in salting meats; so, for example, 60 parts of common salt, 1 part of saltpeter, and 10 parts of sugar, all dissolved in 400 parts of water, make *a good solution for salting beef or pork.*

Potassium chloride, KCl, is a white crystalline salt. It is readily soluble in water and tastes salty. It is chiefly obtained from the Stassfurt deposits and is *the commonest and cheapest of the potassium salts. In crude form it serves as a fertilizer.* From it caustic potash, saltpeter, and other

potassium salts are prepared. This salt is also found in sea water. In the blood and urine of herbivorous animals it is also present in minute amounts.

Potassium bromide, KBr, and **potassium iodide**, KI, crystallize in cubes, which are extremely soluble in water. Both of these salts are used in medicine and in photography.

Potassium chlorate, $KClO_3$, is produced by conducting chlorine into hot solutions of caustic potash, thus:

$$\underset{\text{caustic potash}}{6\,KOH} + \underset{\text{chlorine}}{3\,Cl_2} = \underset{\substack{\text{potassium}\\\text{chlorate}}}{KClO_3} + \underset{\substack{\text{potassium}\\\text{chloride}}}{5\,KCl} + \underset{\text{water}}{3\,H_2O}$$

By electrolyzing hot potassium chloride solutions both caustic potash and chlorine are simultaneously formed, and so most of the potassium chlorate in the market is prepared by this means. *Potassium chlorate is used in making oxygen, matches, fireworks, and explosives.* It is also used in medicine, particularly as a *gargle for sore throat.*

Potassium silicate, K_2SiO_3, is made by fusing sand with potassium carbonate, thus:

$$\underset{\text{silica}}{SiO_2} + \underset{\text{potassium carbonate}}{K_2CO_3} = \underset{\text{potassium silicate}}{K_2SiO_3} + \underset{\text{carbon dioxide}}{CO_2}$$

This salt is soluble in water and its sirupy solutions are popularly called *potassium water glass.*

Potassium sulphate, K_2SO_4, is found at Stassfurt in kainite. It is also a constituent of potassium alum. In crude form *it serves as a fertilizer. It is employed in the manufacture of hard glass, alum, and potassium carbonate.*

Potassium phosphate, K_2HPO_4, is present in minute amounts in the blood and urine of all carnivores. It is a white crystalline salt which dissolves readily in water.

Potassium cyanide, KCN, is an *extremely poisonous* salt. It

is used in photography, in extracting gold from its ores, in gold and silver plating, and also in preparing hydrocyanic acid gas, HCN, for fumigating trees. Potassium cyanide and hydrogen cyanide (also known as **Prussic acid**) are *terrible poisons and must be used with extreme care.*

To the Bunsen flame potassium salts impart a beautiful violet color.

Metallic potassium itself was prepared in 1807 by electrolysis of molten caustic potash by Sir Humphry Davy. Its specific gravity is 0.865 at 15° C. The metal is rather soft, melts at 62°.5 C., and boils at 667° C. With water it reacts violently, forming hydrogen and caustic potash, hence the metal must be kept under petroleum oils to protect it from the moisture of the air.

Metallic sodium and **metallic lithium** have properties that are in general similar to those of metallic potassium. They may be obtained by electrolysis of molten caustic soda, NaOH, and caustic lithia, LiOH, respectively. Sodium and lithium are therefore also kept under petroleum to protect them from the moisture of the air. Sodium is silver-white, somewhat harder than potassium, melts at 95°.5 C. and boils at 742° C. At 15° C. its specific gravity is 0.974.

All sodium salts are soluble in water. To the Bunsen flame they all impart a yellow color, whereas lithium salts color the flame a characteristic red.

In general the *sodium salts are similar to those of potassium. The chief source of all sodium compounds is* **common salt, sodium chloride**, NaCl, which occurs in the sea, in the waters of salt lakes and wells, and also in layers in solid form, particularly at Stassfurt and Reichenhall in Germany, in Galicia, in England, in the states of New York, Michigan, Texas, Utah, and California, also in Africa and Asia. *Salt is very widely distributed on the globe. It is necessary to animal life,*

Sir Humphry Davy, 1778–1829, the discoverer of the alkali metals.

and occurs in the blood, the tissues, and all of the secretions. Just as *land plants take up potassium chloride, so sea plants take up common salt.* Common salt crystallizes in cubes and melts at about 800° C. It is about as soluble in cold as in hot water. At room temperature 100 parts of water dissolve 36 parts of salt, while the same amount of boiling water dissolves but 39 parts. *The United States alone produces nearly four and one-half million tons of salt annually* and this is only about one-fourth of the world's yearly production. Salt is used for making chlorine, hydrochloric acid, caustic soda, sodium carbonate, and other sodium compounds.

Caustic soda, sodium hydroxide, NaOH, is quite similar to potasium hydroxide in its properties and is prepared in an analogous manner. It is used in making hard soap, also in " softening " water. It is further employed in the manufacture of carbolic acid and oxalic acid, also in making paper and other articles. *Being cheaper than caustic potash, it is used in place of the latter whenever it is possible.*

Sodium carbonate, Na_2CO_3, is also often called **sal soda**, or **soda**. Just as potassium carbonate or potash occurs in the ashes of land plants, so *soda is found in the ashes of marine plants. Soda is used to make common glass, hard soap, and many other sodium compounds.* It is therefore manufactured on a very large scale. *There are two processes of manufacturing soda, namely, the LeBlanc process and the Solvay process.* Both utilize common salt as the raw material.

In the **LeBlanc soda process** salt is first treated with sulphuric acid, thus:

$$\underset{\text{salt}}{2\,NaCl} + \underset{\text{sulphuric acid}}{H_2SO_4} = \underset{\text{salt cake}}{NaCl + NaHSO_4} + \underset{\text{hydrochloric acid}}{HCl}$$

The so-called " salt cake " consists of salt and acid sodium sulphate. This is then heated in a suitable fur-

nace, thus forming more hydrochloric acid, and sodium sulphate:

$$\underset{\text{salt cake}}{NaCl} + NaHSO_4 = \underset{\substack{\text{sodium}\\\text{sulphate}}}{Na_2SO_4} + \underset{\substack{\text{hydrochloric}\\\text{acid}}}{HCl}$$

The sodium sulphate is then mixed with charcoal and calcium carbonate and heated in a rotary furnace; in this way carbon monoxide escapes and calcium sulphide and soda are formed, thus:

$$\underset{\text{sodium sulphate}}{Na_2SO_4} + \underset{\text{carbon}}{4\,C} = \underset{\text{sodium sulphide}}{Na_2S} + \underset{\text{carbon monoxide}}{4\,CO}$$

and

$$\underset{\substack{\text{sodium}\\\text{sulphide}}}{Na_2S} + \underset{\substack{\text{calcium}\\\text{carbonate}}}{CaCO_3} = \underset{\text{soda}}{Na_2CO_3} + \underset{\substack{\text{calcium}\\\text{sulphide}}}{CaS}$$

The entire mass is then treated with water. The sodium carbonate dissolves and calcium sulphide remains as an insoluble residue. From the clear solution, crystals are obtained which have the composition $Na_2CO_3 . 10\,H_2O$. This is so-called **washing soda** or **crystallized soda**. On heating the crystals they lose water and form **calcined soda** or *anhydrous sodium carbonate*, Na_2CO_3.

By the **Solvay process**, sodium bicarbonate, $NaHCO_3$, is first obtained. This salt is relatively less soluble than the corresponding ammonium salt, **ammonium bicarbonate**, NH_4HCO_3, which may be obtained when carbon dioxide is passed into strong ammonia water, thus:

$$\underset{\substack{\text{ammonium}\\\text{hydroxide}}}{NH_4OH} + \underset{\substack{\text{carbon}\\\text{dioxide}}}{CO_2} = \underset{\substack{\text{ammonium}\\\text{bicarbonate}}}{NH_4HCO_3}$$

On treatment of a saturated solution of sodium chloride with ammonium bicarbonate, sodium bicarbonate is precipitated, thus:

$$\underset{\text{salt}}{NaCl} + \underset{\substack{\text{ammonium}\\\text{bicarbonate}}}{NH_4HCO_3} = \underset{\substack{\text{sodium}\\\text{bicarbonate}}}{NaHCO_3} + \underset{\substack{\text{ammonium}\\\text{chloride}}}{NH_4Cl}$$

The ammonium chloride remains in solution. In actual practice, then, *the Solvay process consists in conducting both ammonia and carbon dioxide into a saturated solution of common salt.* Thus both of the changes expressed by the last two equations take place together. Sodium bicarbonate is also called **saleratus** or **baking soda.** It has a mild alkaline

FIG. 39. — A Chilian saltpeter bed.

reaction, and is used in medicine, in baking powders, etc. On heating sodium bicarbonate it loses water and carbon dioxide and forms sodium carbonate, thus:

$$2\ NaHCO_3 = Na_2CO_3 + H_2O + CO_2$$

Even when solutions of sodium bicarbonate are heated, carbon dioxide is evolved; *upon this fact the use of saleratus in baking depends.*

Sodium nitrate, $NaNO_3$, is found in large beds in Chili and is consequently known as **Chili saltpeter.** With it occur

smaller amounts of sodium chloride, sodium sulphate, and **sodium iodate**, $NaIO_3$. The latter compound is *the chief commercial source of iodine.* Sodium nitrate is used as a fertilizer on account of its nitrogen content. It also is employed in large quantities in making nitric acid, thus:

$$NaNO_3 + H_2SO_4 = NaHSO_4 + HNO_3$$

Sodium nitrate should be added to the soil only in small amounts at a time.

Sodium sulphate, Na_2SO_4, is formed as one of the products of the Leblanc soda process. It dissolves readily in water, and when it crystallizes from the latter at room temperatures beautiful crystals of the composition $Na_2SO_4 . 10\ H_2O$ are formed. This is the so-called **Glauber's salt**, which is so much *used as a laxative.* It was first made by Johann Rudolf Glauber in 1658.

Sodium sulphite, Na_2SO_3, is made by passing sulphur dioxide into sodium hydroxide solution. This salt is used in photography.

Sodium thiosulphate, $Na_2S_2O_3 . 5\ H_2O$, is the so-called "**hyposulphite of soda**" or "**hypo**," so much used in making the "fixing bath" in photography. It is produced by boiling a solution of sodium sulphite with flowers of sulphur.

Sodium silicate, Na_2SiO_3, or sodium **water glass**, is produced by fusing silica with sodium carbonate or with a mixture of carbon and Glauber's salt. In laundry soaps it serves as a "filler." It is also employed in cementing together the fibers of mineral wool and asbestos, and in fireproofing wood, fabrics, etc.

Borax has the composition $Na_2B_4O_7 . 10\ H_2O$. Its uses have already been mentioned.

The **metals of the alkaline earths** are barium, strontium, calcium, and magnesium. They are so called because their respective oxides BaO, SrO, CaO, and MgO are alkaline,

which fact is apparent when they are treated with moist red litmus paper, for example. These metals may be prepared by the electrolysis of their molten chlorides. They all are silver-white in appearance, and when treated with water they yield hydrogen and the hydroxide of the metal employed. The hydroxides are soluble in water, but not copiously.

The chief source of barium is **barium sulphate**, $BaSO_4$, also called **heavy spar** or "**barytes.**" It occurs in nature in fairly pure form in compact masses. Its specific gravity is 4.48. Ground to powder, it serves as a white pigment in paints, being called "**permanent white.**" It is not infrequently used as *a cheap adulterant in white lead,* a practice that is to be condemned. Barium sulphate is almost completely insoluble in water and also in dilute acids. It precipitates whenever a solution of a sulphate is treated with barium chloride solution or a solution of any other barium salt, so, for instance:

$CuSO_4$	+	$BaCl_2$	=	$BaSO_4$	+	$CuCl_2$
copper sulphate		barium chloride		barium sulphate		copper chloride

Barium chloride is consequently very commonly used in testing for sulphates, also for determining the amount of sulphur present in a compound, for the latter process commonly consists of converting the sulphur into a soluble sulphate and then precipitating barium sulphate by adding an excess of barium chloride solution. From the amount of barium sulphate formed the amount of sulphur may be computed, for the composition of barium sulphate is well known. **Barium nitrate,** $Ba(NO_3)_2$, is used for making *green Bengal lights.* Barium salts are all poisonous. Cattle in our Western states have been reported as seriously poisoned by eating herbs that had taken up barium from the soil. This

is, of course, very unusual, for barium is ordinarily not found in soils.

Strontium compounds are entirely similar to those of barium, only they are of rarer occurrence. **Strontium nitrate**, $Sr(NO_3)_2$, is used in fireworks and for making *red Bengal lights*. All strontium salts color the Bunsen flame red, while those of barium color it green, and calcium compounds produce an orange-yellow color.

Calcium carbonate, $CaCO_3$, is found in large quantities in nature as *limestone, marble, and chalk*. Iceland spar or calcite is a very pure crystalline form of calcium carbonate. *When limestone occurs mixed with clay, it is called marl.* Calcium sulphate, $CaSO_4$, occurs as anhydrite, **gypsum**, $CaSO_4 . 2\ H_2O$, and alabaster. *In natural waters calcium bicarbonate and calcium sulphate are quite commonly found.* They produce the hard sediments in teakettles, boilers, etc. These sediments generally also contain silica, oxides of iron, aluminum, magnesium, etc. These have been dissolved from the rocks or soils with which the waters have come in contact.

Marble and **limestone** serve as building stones, also for making lime, glass, and cement, and reducing iron ores. In the form of **chalk** calcium carbonate is used as **whiting**, which when ground up with linseed oil forms **putty**.

Lime, CaO, is made by heating limestone or marble in kilns. Thus, carbon dioxide is driven off and the oxide of calcium remains:

$$\underset{\text{limestone}}{CaCO_3 \text{ on heating}} = \underset{\text{lime}}{CaO} + \underset{\text{carbon dioxide}}{CO_2}$$

The slaking of lime and its use in making mortar have already been described.

Calcium chloride, $CaCl_2$, is formed when hydrochloric acid acts on limestone, thus:

$$\underset{\text{limestone}}{CaCO_3} + \underset{\text{hydrochloric acid}}{2\,HCl} = \underset{\text{calcium chloride}}{CaCl_2} + \underset{\text{water}}{H_2O} + \underset{\text{carbon dioxide}}{CO_2}$$

Calcium chloride is very soluble in water, and its brines are commonly used as the cold liquid that circulates in the pipes in refrigerators. This brine is kept cold by means of evaporation of ammonia which has been liquefied under pressure.

FIG. 40. — A modern lime kiln.

On carefully heating gypsum, $CaSO_4 . 2\,H_2O$, to about 110° it loses water and becomes $CaSO_4 . \frac{1}{2}\,H_2O$, which is known as **plaster of Paris**. When it is afterward mixed with water, the whole sets or hardens to a solid mass, which depends upon the fact that plaster of Paris and water unite and again form gypsum, thus:

$$\underset{\text{plaster of Paris}}{CaSO_4 . \tfrac{1}{2}\,H_2O} + \underset{\text{water}}{1\tfrac{1}{2}\,H_2O} = \underset{\text{gypsum}}{CaSO_4 . 2\,H_2O}$$

Plaster of Paris is much used in making hard wall plaster, gypsum, calsomine, casts, and bandages in surgery.

As all plants and animals contain sulphur and calcium salts, calcium sulphate in the form of gypsum is commonly used as a fertilizer under the name of **land plaster.**

FIG. 41. — A broken leg in a plaster of Paris cast.

Calcium nitrate, $Ca(NO_3)_2$, is very soluble in water. It is formed in composted manure, and also on the white-washed walls of stables. The ammonia present in the stable air is gradually oxidized to nitric acid, which then attacks the calcium carbonate on the walls, forming calcium nitrate and carbon dioxide. This, of course, destroys the wall coating. The *calcium nitrate in the soil is a good plant food, for it contains both calcium and nitrogen in available form.*

Calcium phosphate, $Ca_3(PO_4)_2$, is insoluble in water but soluble in dilute acids. *It occurs in crystals as apatite, is the chief constituent of bones, and is found in phosphate rock, guano, and barnyard manure.* The fact that by treating bones or phosphate rock with sulphuric acid, superphosphate fertilizer is formed has already been mentioned. *Over two and one-half million tons of phosphate rock are used annually in the United States as fertilizer.* The exportation of phosphate rock should be forbidden by law, for the supply of this important material, without which agriculture cannot thrive, is limited. The **beds of phosphate**

rock occur mainly in Florida and South Carolina, but recently deposits have also been found in some of the Rocky Mountain states.

As **calcium silicate**, calcium occurs in many rocks, especially in complex silicates like feldspars, garnet, mica, hornblende, etc.

Portland cement is made by heating limestone together with aluminum silicates in a kiln, and then grinding up the clinker to a fine powder. This is usually grayish or grayish

FIG. 42. — Concrete fence posts.

brown in color. When mixed with water it will unite with the latter, forming a hard stonelike mass which in appearance resembles the native stone found at Portland, England, whence the name Portland cement. No chemical formula can be given for Portland cement, but a good Portland cement must contain silica, SiO_2, from 20 to 25 per cent; lime, CaO, from 58 to 65 per cent; alumina, from 5 to 10 per cent. Magnesia, MgO, must not exceed 5 per cent, and the less

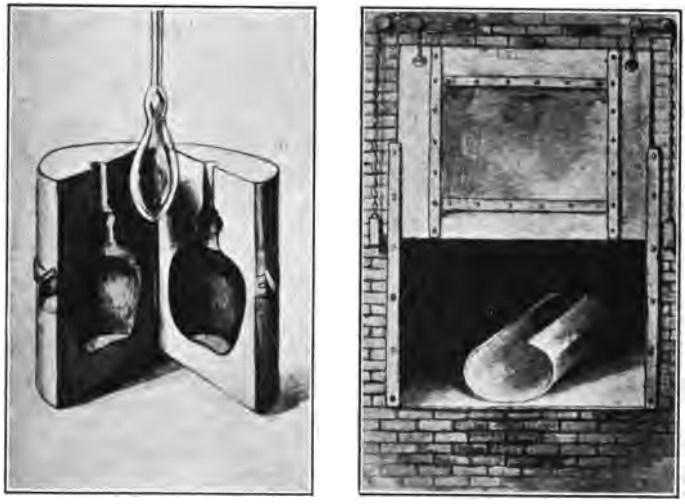
there is of it, the better. Iron oxide, Fe_2O_3, may be present in notable quantities without doing harm; it is commonly found in cements to the extent of from 2 to 5 per cent. Other impurities, like small amounts of soda, potash, and sulphates, are generally also found, commonly less than 1 per cent.

At Milwaukee, Louisville, and Rosendale, N.Y., occur natural deposits of limestone that contain almost the right amounts of aluminum silicates to make Portland cement. This material is heated in kilns, and the resulting clinker is then ground up and sold as "**natural cement.**" While quite serviceable for many purposes and cheaper than Portland cement, it nevertheless is not equal to the latter in wearing qualities. *Portland cement sets, even under water,* and hence can be used in damp places where ordinary lime mortar will not harden at all. Mixed with sand and crushed stone or gravel and then saturated with water, Portland cement forms so-called **concrete.** In modern structures, iron or steel rods or networks of wires are often embedded in concrete to strengthen it. This is then known as **reënforced concrete.** It is much used in erecting bridges, fireproof buildings, etc. *Nearly fifteen thousand tons of Portland cement are produced annually in the United States alone.* The slags from blast furnaces, which contain silica, alumina, and lime, are now also used in the manufacture of Portland cement, the lacking ingredients being added, of course, before firing the material in the kiln. The cements from blast furnace slag are generally richer in iron and browner in color than other cements. They are of very good quality, nevertheless.

Glass commonly consists of a mixture of the silicates of calcium and sodium. *It is made by melting together silica,* SiO_2, *limestone,* $CaCO_3$, *and soda,* Na_2CO_3. The final product corresponds approximately to the composition

$$CaO \cdot 3\,SiO_2 + Na_2O \cdot 3\,SiO_2$$

and is known as **soft glass** or *soda lime glass*. It is used in making windows, bottles, and ordinary glass dishes. The mixture is melted in pots of fire clay about four feet high and four feet in diameter. *Potash lime glass*



(*A*) (*B*)

FIG. 43. — (*A*) Blowing a glass bottle. (*B*) Making window glass.

is **hard glass**. It is also called *Bohemian glass*. It contains potassium silicate instead of sodium silicate. It is not as readily attacked by water, acids, and alkalies as soda lime glass, and hence is used for making glassware for chemical purposes, like beakers, retorts, etc. **Crown glass** is also a potash lime glass. **Jena glass** contains boric anhydride. *Blue glass* is colored by cobalt oxide; *brown glass* by ferric oxide; *green glass* by ferrous oxide or chromic oxide; etc. *Black glass* contains large amounts of iron and other metallic oxides. **Milk glass** contains suspended particles of calcium

phosphate. The Egyptians and Phœnicians knew how to make glass long before the Christian era. Window glass was first used in the sixteenth century.

Metallic magnesium is sold in the market in powder, and in the form of wire or ribbon. It burns with a brilliant white light, forming magnesium oxide, MgO. The metal is used *for fireworks, for flash lights in signalling and in photography.*

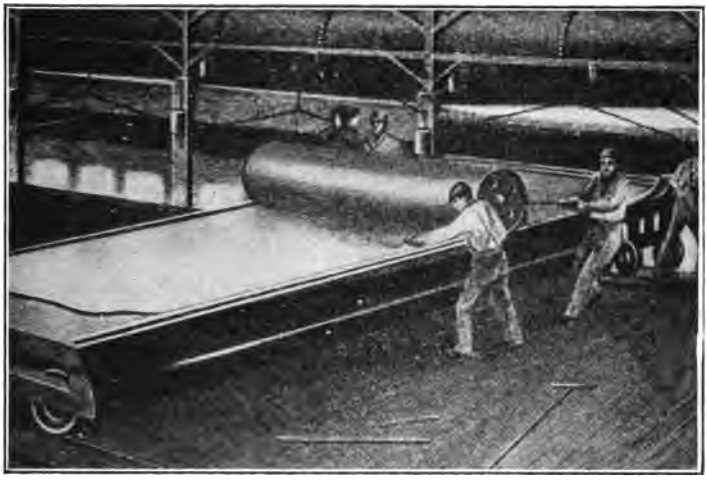


FIG. 44. — Manufacture of glass. Rolling out plate glass.

Five parts of magnesium powder to nine parts of pulverized potassium chlorate makes a good **flash-light powder.**

In nature magnesium is found as *magnesite,* the carbonate $MgCO_3$, but far more commonly as **dolomite** or **magnesium limestone,** $MgCO_3 . CaCO_3$, which has already received mention. Magnesium occurs in many natural silicates very widely distributed. So there are **hornblende** $Mg_2CaFeSi_4O_{12}$, **asbestos** $Mg_3Si_2O_7 . 2\ H_2O$, **meerschaum** $Mg_2Si_3O_8 . 4\ H_2O$, and many other complex silicates that contain magnesium. All plants and animals contain magnesium compounds, and

these finally appear in the ash as phosphates and carbonates. *Magnesium ammonium phosphate occurs in guano and is valuable as a fertilizer.*

Magnesium oxide and **magnesium carbonate** are used in medicine. The latter compound also serves as a *face powder.* **Talc,** a hydrous magnesium silicate, is used for the same purpose. **Magnesium sulphate** serves as a purgative under the name **Epsom salt.** *All magnesium salts have a bitter taste and act as laxatives.* When found in natural waters, the latter are called *bitter waters.* Some of them, like the waters of the springs at Epsom in England and Hunyad in Hungary, are quite famous. It is to be borne in mind that while the carbonates and phosphates of barium, strontium, calcium, and magnesium are *difficultly soluble in water,* and the sulphate of barium is *insoluble,* the sulphates of strontium and calcium are *sparingly soluble,* the sulphate of magnesium is quite *readily soluble.* Four parts of water dissolve one part of magnesium sulphate crystals at room temperatures.

In their chemical behavior, magnesium compounds are in many ways analogous to those of zinc.

Magnesium chloride, which is cheap, as it is a by-product of the potash salt industry at Stassfurt, forms with magnesium oxide a basic magnesium chloride. This sets as a hard mass, and is frequently used as *a cement in making floors, wainscoting, etc.* The oxide obtained by calcining native magnesium carbonate is mixed dry with the appropriate amount of sand, sawdust, or other filler and the desired pigment, and then this powder is treated with magnesium chloride solution of proper strength, as in making a mortar. The material is spread out with trowels, and sets in about 24 to 30 hours to a hard mass of good wearing qualities. In Germany the material is called "**Steinholz.**" In America this cement is being intro-

duced, and various fanciful trade names are being applied to it.

The salts of the element **radium** are similar to those of the metals of the alkaline earths. *Radium salts glow in the dark.* The light affects photographic plates. The intensity of this light increases with the purity of the radium compounds. *Heat is also constantly evolved by radium compounds, and the air in their immediate neighborhood has less electrical resistance than ordinary air.* This is due to the radiations or so-called emanations that are continually being emitted by radium compounds. In extremely minute amounts, which indeed are almost infinitesimal, radium compounds are very widely distributed in soils and natural waters. In uranium compounds, especially in an oxide of uranium known as **pitchblende**, radium occurs in *relatively* larger quantities. *The emanations from radium are able to destroy germs. They also stop the germinating power of seeds and kill the tissues of living beings.* Experiments with radium compounds must therefore be conducted with great care. Their value as an agent for curing diseases is still a question to be determined.

QUESTIONS

1. Name the alkali metals and state their general characteristics.
2. What is potash and how is caustic potash prepared from it?
3. Where is potassium found in nature?
4. Mention five important compounds of potassium.
5. What is black gunpowder chemically?
6. How is potassium chlorate made, and what is it used for?
7. What is potassium cyanide? What are its properties and uses?
8. How may sodium compounds be distinguished from potassium compounds?
9. Where is common salt found in nature? What is salt used for?

10. What are the two processes for making soda on a commercial scale? What uses are made of soda?

11. What is the difference between washing soda and baking soda?

12. (*a*) What is Glauber's salt? (*b*) What is the "hypo" used in photography? (*c*) What is sodium water glass, and for what is it used?

13. Name the metals of the alkaline earths and state why they are so called.

14. What is barium sulphate? What is it used for? By what other names is it known?

15. How much barium is there in a ton of barium sulphate?

16. What use is made of strontium nitrate?

17. What are the various forms in which calcium carbonate occurs? How do we know that these are all one and the same chemically?

18. What is gypsum? Plaster of Paris?

19. What happens when plaster of Paris sets?

20. Explain the action of hydrochloric acid on limestone.

21. Of what value is calcium nitrate in soils?

22. What is phosphate rock and why is it valuable?

23. How is Portland cement made?

24. What is "natural cement"? Where is it made?

25. What are the uses of Portland cement?

26. What is common glass chemically? How does soft glass differ chemically from hard glass? How is colored glass produced? When was window glass first made?

27. For what purposes is metallic magnesium used?

28. Mention five compounds of magnesium that occur in nature.

29. What is Epsom salt chemically?

30. Mention some of the characteristics of compounds of radium.

CHAPTER XI

ALUMINUM, THE HEAVY METALS, AND THEIR IMPORTANT ALLOYS

Of all the metals aluminum is by far the most abundant, forming about 8 per cent of the material of which the earth's crust is composed. It is found mainly as a constituent of all the common siliceous rocks, especially in feldspars, clays, micas, granites, slates, etc. *It never occurs in the uncombined state,* for it has great affinity for oxygen, and indeed it nearly always is found in compounds with that element. **Emery** is an impure form of aluminum oxide, being colored brown by the presence of oxides of iron. **Sapphires** and **rubies** are beautifully crystallized aluminum oxide and are highly prized as gems. **Bauxite** is a hydrated oxide of aluminum, and *is used as a source for the manufacture of alumi-*

Fig. 45. — An emery wheel.

num. In Greenland a double fluoride of sodium and aluminum, $AlF_3 . 3\ NaF$, is found. It forms white masses insoluble in water. This mineral is called **cryolite**. It may be melted quite readily and *the molten mass dissolves aluminum oxide.* By passing the electric current through this molten mass metallic aluminum is deposited and oxygen comes off at the other pole. The bath is replenished by continually feeding in more aluminum oxide, also called **alumina**. The molten mass is contained in a vessel of graphitic carbon which serves

FIG. 46. — Cooking utensils made of aluminum.

as the pole on which the metal is deposited. Carbon sticks dipped into the molten mass form the other pole. Eight volts electrical pressure is sufficient to keep the process going, but the current strength is usually several thousand amperes. Thus *the heat developed by the electric current is itself sufficient to keep the bath and the deposited metal as well in the molten*

state, once the process has been started. The molten aluminum is tapped off from the bottom of the vessel from time to time. *Over twenty-five thousand tons of aluminum are thus produced annually in the United States alone.* The metal sells for less than twenty cents per pound. **Aluminum metal** is silver-white in appearance, though on exposure to the air it oxidizes slowly, and the film of oxide on its surface gives the metal a bluish appearance. Aluminum is only about one-third as heavy as iron. It is malleable, ductile, melts at 660°, and is *a good conductor of electricity and also of heat.* It is used for wires and electric cables, for cooking utensils, and many other useful articles.

Magnalium is an alloy consisting of 75 to 90 per cent of aluminum and 25 to 10 per cent magnesium. It is harder and lighter than aluminum and may be polished to a higher degree. **Aluminum bronze** consists of 90 to 95 per cent copper and 10 to 5 per cent aluminum. It is yellow in color, hard and strong, and is consequently often used in the arts.

Aluminum paint consists of finely divided aluminum suspended in a suitable varnish or laquer.

Mixed with the black oxide of iron, Fe_3O_4, finely divided metallic aluminum forms a mixture called **thermite.** When once this mixture is ignited (which can be accomplished by means of a fuse of magnesium ribbon to which is attached a mixture of magnesium powder and potassium chlorate) the chemical action continues rapidly and vigorously, the temperature developed being about 3000° C. The change that takes place is that aluminum robs the oxide of iron of its oxygen and so forms alumina and metallic iron, thus:

$$\underset{\text{black iron oxide}}{3\ Fe_3O_4} + \underset{\text{aluminum}}{8\ Al} = \underset{\text{alumina}}{4\ Al_2O_3} + \underset{\text{iron}}{9\ Fe}$$

Thermite is used for welding car rails and making welds in

many parts of machinery of various kinds. The parts need only to be butted together, a mold built around the joint, and the molten thermite mixture run upon it. Thus the iron is heated to the welding point very quickly. In many cases

FIG. 47. — Welding a car rail with thermite.

it is not even necessary to take the machinery apart to make the repairs, the process is so simple.

Aluminum sulphate, $Al_2(SO_4)_3$, is made by dissolving aluminum hydroxide, $Al(OH)_3$, in sulphuric acid, thus:

$$2\ Al(OH)_3 + 3\ H_2SO_4 = Al_2(SO_4)_3 + 6\ H_2O$$

This salt is readily soluble in water and is used as a mordant for fixing dyestuffs upon fabrics. In the paper industry it serves in sizing paper.

Ordinary **potassium alum** is a double sulphate of aluminum and potassium. It has the composition $K_2SO_4 . Al_2(SO_4)_3 . 24\ H_2O$. **Ammonium alum** and **sodium alum** contain ammonium and sodium respectively instead of potassium, otherwise they are quite like potassium alum. Alum crystal-

lizes in octahedra, dissolves readily in water, and like all other soluble aluminum salts, it has an astringent effect on the mucous membranes. *Alum is used as a mouth wash in medicine, as a mordant in dyeing fabrics, and unfortunately it is also still used in baking powders.* **Alum baking powders** contain alum or aluminum sulphate and baking soda. When this mixture is moistened, as, for example, in baking, carbon dioxide is evolved which raises the dough. There is simultaneously formed sodium sulphate and aluminum hydroxide, and these remain in the biscuit, bread, or cake. The aluminum hydroxide is readily dissolved by the hydrochloric acid in the stomach, thus forming aluminum chloride, a powerful astringent which interferes with digestion and is harmful to the linings of the digestive tract. Bread and cake made with alum baking powder looks well, for the carbon dioxide is liberated steadily and just about at the right rate to make fine-appearing cookery. *As alum baking powders are cheap they are still unfortunately much in use,* even though in some states the law requires the fact that they contain alum or aluminum sulphate to be stated on the label to warn the public.

It is a notable fact that though aluminum silicates are present in every soil, plants and animals only contain minute traces of aluminum.

Clay consists essentially of aluminum silicates that have been formed from the weathering of feldspars and other complex silicates that make up the granitic rocks. In its purest form clay is white and is called **kaolin**, $H_2Al_2(SiO_4)_2 . 2\,H_2O$. This is used for making **white porcelain ware.** *Common clay is discolored by impurities.* Usually it is brown or red, which is caused by the presence of oxides of iron. Ordinary clay serves for making bricks, flowerpots, and the inferior and cheaper grades of pottery and crockery ware.

In making **bricks** and porous **earthenware** the clay is molded into the desired shapes and then heated to redness in kilns, or "fired" as it is called. *A cheap glaze may be put upon such ware by simply throwing salt into the kiln.* Thus as the water is baked out of the clay, the water vapor at the high temperature that obtains decomposes the salt, forming hydrochloric acid and sodium hydroxide. The latter unites with the surface of the clay, forming an easily fusible silicate which

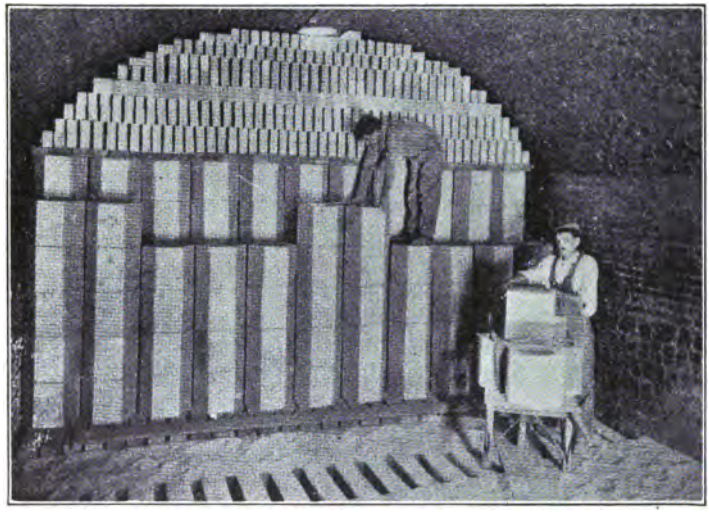


FIG. 48. — A pottery kiln. Setting the shapes and getting ready to fire.

produces a thin glassy coating or so-called *glaze. Butter jars and other "stoneware" crocks are glazed in this way.* **Porcelain** is vitreous throughout. This is accomplished by mixing finely pulverized feldspar and quartz with the kaolin before molding the dishes. When these are dried and finally fired, the feldspar and quartz fuse and fill the pores of the ware so that it exhibits a perfectly vitreous fracture instead of an earthy one, as the cheaper earthenware, either glazed or

unglazed, always does. **Fire bricks** contain a larger amount of silica than ordinary ones and are consequently more refractory. **Fire clay** also is especially rich in silica. **Colored porcelain** and **colored glazes** are produced by means of various metallic oxides, as in making colored glass, for example.

Ultramarine is a fine dark blue pigment made by heating together clay, soda, sulphur, and charcoal out of contact with the air. It is probably a double compound of sodium aluminum silicate and sodium sulphide. Ultramarine is attacked by acids, which destroy its color, but in presence of alkalies it is quite stable. *It serves as laundry blue, also as a pigment in paints, and in the manufacture of wall paper.* Furthermore, *it is used to destroy the yellow appearance of sugar, linen, cotton, and paper pulp,* only enough being used to produce a pure white appearance.

Aluminum oxide, being a white powder that is practically insoluble in water and neutral in reaction toward litmus, is a typical " earthy " oxide, and aluminum is therefore termed an *earth* metal. There are several other similar **earth metals,** but they are so very rare and have no importance in practice that they need not be considered here. The oxides of cerium and thorium belong to these **rare earths.** It has already been stated that these are used in making mantles for the Welsbach lamps. In the **Nernst lamp** a filament of earthy oxides is heated to incandescence by passing the electric current through it.

Platinum, gold, and silver are called the noble metals. They occur in nature in the free or uncombined state and are not at all readily oxidized.

Platinum is silver-white, very malleable, and ductile. It occurs in alluvial sands in the Ural Mountains, in California, Brazil, and Australia. It is *not* attacked by hydrochloric, nitric, or sulphuric acid, nor by molten sodium or potassium

carbonate, though **aqua regia**, a mixture of nitric and hydrochloric acid, will dissolve it, forming platinic chloride. Platinum does *not* melt readily, its melting point being about 1777°. These facts make it very valuable as a material out of which to construct dishes for certain kinds of chemical work in which glass and porcelain would be attacked. *Platinum is also used for electrical contacts and for spark points in the spark plugs of gasoline engines.* For the latter purpose it is commonly *alloyed with iridium*, a metal which is similar to platinum. This alloy is the so-called **hard platinum**. Platinum salts are used in photography, and also in chemistry for the purpose of estimating potassium, for **potassium platinic chloride**, K_2PtCl_6, is a beautiful, crystalline, yellow salt which is difficultly soluble in dilute alcohol. It is one of the very few potassium salts that do not dissolve copiously in water. *In a finely divided form platinum absorbs oxygen which is readily given off to oxidizable substances.* So, for example, one can light a gas jet by bringing it into contact with such finely divided platinum, also called **platinum sponge**. Other metals of the platinum family are **iridium**, **osmium**, **ruthenium**, **rhodium**, and **palladium**. They all occur with platinum in nature and exhibit somewhat similar properties. *The many uses to which platinum has been put and the small supply make the metal very costly.* It is worth about forty-five dollars per ounce.

FIG. 49. — A modern spark plug with platinum-iridium points.

Gold, aurum, has always been regarded as an article of value. It melts at 1064°, *is extremely malleable and ductile.* In aqua regia it dissolves, forming gold chloride. **Gold coins**

contain 9 parts gold and 1 part copper. This alloy is much harder than pure gold, which is rather soft and consequently will not wear well. For the same reason **copper alloys of gold** are used for **jewelry**, ornaments, etc. Pure gold is 24 carats fine; 16 carat gold contains 16 parts gold and 8 parts copper; 14 carat gold contains 14 parts gold and 10 parts copper; etc. The United States produces about 160 tons of gold annually, valued at about $96,500,000. The whole world produces about four times this amount per year. Gold salts are used in photography as "toning baths." *Many metallic articles are plated with gold.* This is accomplished by immersing the article to be plated in a bath consisting of a solution of gold potassium cyanide, $KAu(CN)_4$, together with an electrode of pure gold; a current of electricity is then passed from the latter electrode through the solution to the article to be plated, which is readily coated with the metal. *Gold is also used on the edges and backs of books, for ornamenting chinaware, gilding signs, church spires, etc.*

Silver, argentum, occurs in nature in the free state, but also as the sulphide and chloride. *It is much more abundant than platinum and gold.* Nearly 2000 tons of silver are produced annually in the United States. The metal melts at 962°, is malleable, ductile, and *the best conductor of electricity and heat known.* **Sterling silver** and **silver coins** contain 1 part copper and 9 parts silver. Nitric acid dissolves silver readily, forming **silver nitrate**, $AgNO_3$, which is *the most important of the silver compounds*, for from it all others are prepared. It is used in medicine and in photography. *It is one of the most soluble of all salts.* Even in ice-cold water it is possible to dissolve 122 parts in 100 parts of water, while in boiling water the solubility increases tenfold. This salt is also called **lunar caustic** and is used for *cauterizing wounds, removing warts, etc.* **Silver chloride**, AgCl, **silver bromide**, AgBr, and

silver iodide, AgI, are all insoluble in water. *These salts darken on exposure to light*, probably because of separation of finely divided silver. *This fact is the basis of the use of these salts in photography.* The **photographic plate** consists essen-

(*A*) (*B*)

FIG. 50. — Photography. (*A*) A positive. (*B*) A negative.

tially of silver bromide embedded in gelatine, which covers one side of the glass. On exposure to light in the camera, the bromide is slightly, but not visibly, reduced. When the plate is afterwards treated with a **developer** (a reducing agent like pyrogallic acid, hydrochinone, etc.), the reduction continues where the light has started it, and so the picture becomes visible. *The developer must be washed off when the*

picture has been sharply developed or the process will continue and produce a blurred outline. To remove the undecomposed silver bromide in the gelatine, the plate (before exposure to light) must be treated with a solution of sodium thiosulphate, $Na_2S_2O_3$, the "**hypo**" **bath**. This dissolves the silver bromide, leaving only the gelatine with the metallic silver particles which form the outline of the picture. The plate is then thoroughly rinsed off and dried. It is a **negative**, *i.e.* it is dark where the object which was photographed was light, and *vice versa*. To make a **positive** this negative is laid upon a bromide **print paper**, which is paper coated with a sensitive film of silver bromide similar to that on the bromide plate. The whole is now exposed to the light. The print must be "fixed" in the "hypo" bath the same as the negative, and it may then be exposed to the light.

Silver plating is done in a solution of potassium silver cyanide, $KAg(CN)_2$. *The process is similar in all respects to gold plating*, which has already been mentioned. The *black stains that form on silverware* (especially on spoons, forks, etc., which have been in contact with eggs) are *silver sulphide*, and not silver oxide as is often thought.

Copper, cuprum, occurs in large quantities in the free state near Lake Superior. In Montana it is found in ores in combination with iron and sulphur. *Nearly half of the annual output of copper in the world is produced in the United States, which contributes about 540,000 tons per year.* Copper and gold are the only colored metals known. Copper melts at 1084°. It is tough, rather hard, ductile, and malleable. *It conducts heat and electricity well.* Much *copper wire is used in electrical work.* For other purposes, however, *copper is generally used in the form of alloys.* These are made by melting the metals together in the desired proportions. The following are some of the most important **alloys of**

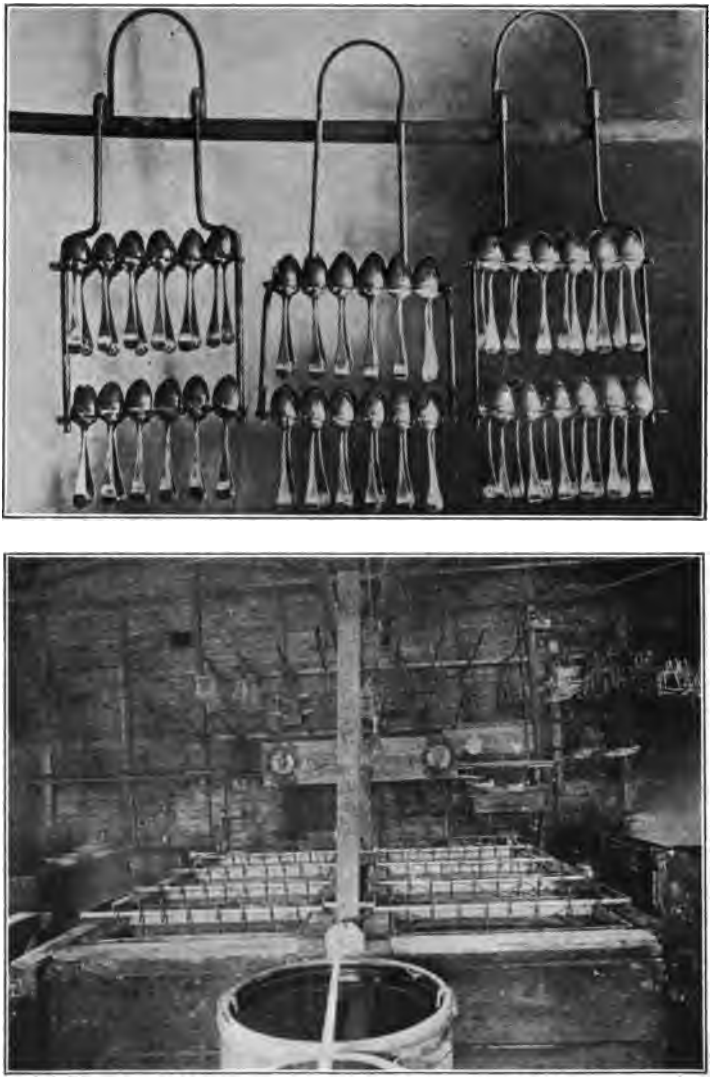


FIG. 51. — Plating silver spoons. The upper figure shows the rack on which the spoons are hung.

copper: **Brass**, which is yellow in color, contains 1 part zinc and 2 parts copper, though other proportions are often used. **Dutch metal**, which is reddish brown, consists of 1 part zinc and 5 parts copper. **German silver** consists of 80 to 95 per cent brass and 5 to 10 per cent nickel. **Gun metal** contains 9 parts copper in 1 part tin. **Bell metal** consists of 3 parts copper plus 1 part tin. *The alloys of copper and tin are called bronzes.* **Bronzes** for statuary commonly contain 3 to 8 parts tin, 1 to 3 parts lead, 1 to 10 parts zinc, and 80 to 90 parts copper. **Phosphor bronze** is used in making parts of machinery. It is especially hard. It consists of bronze to which from 0.5 to 3 per cent phosphorus has been added. *Copper dissolves readily in nitric acid,* also in hot concentrated sulphuric acid, but cold sulphuric or hydrochloric acid has but very little effect on it. However, in presence of the air, many dilute acids do very gradually attack copper, which fact must be kept in mind, for copper and brass are frequently used for cooking utensils, and *copper compounds are poisonous.* On copper roofs, old coins, etc., that have been exposed to moist air for a long time there is formed a green deposit called **verdigris**. It is a basic carbonate of copper, $CuCO_3 . Cu(OH)_2$. *In ammonia water copper is soluble when in contact with the oxygen of the air.* However, ammonia acts but slowly, hence it is often used for **cleaning copper**. *The most important as well as the most common salt of copper is copper sulphate, blue vitriol,* $CuSO_4 . 5\ H_2O$. It forms large blue crystals, is soluble in about 3 parts of water, and its solutions are used as a bath for **copper plating**, for spraying plants, especially in making Bordeaux mixture, and also for preparing other copper compounds. In Paris green copper also is present, as has been mentioned before.

Mercury, quicksilver, hydrargyrum, is the only metal which is a liquid at ordinary temperatures. It is found in

nature in the free state, but generally it is combined with **cinnabar**, HgS, which is red in color and is used as a pigment, being called **vermilion**. Mercury comes from Spain, Austria, Prussia, California, Japan, and China. *It is used in thermometers, barometers, in amalgams for the backs of mirrors, also for filling teeth. Its compounds are used in medicine.* So mercurous chloride, HgCl, is **calomel**. It is but slightly soluble in water. Mercuric chloride, $HgCl_2$, is **corrosive sublimate**. It is very copiously soluble in water and is an *exceedingly powerful poison.* It is to be borne in mind, moreover, that *all mercury compounds are poisonous. The antidote is raw eggs and milk.* The albumin forms an insoluble compound with the mercury salt, which is then got rid of by an emetic or a purge. Mercury melts at $-39°.4$ C. and boils at 357° C. Solid mercury may be hammered into sheets and cut with tools like other metals. The **alloys of mercury** with other metals are called **amalgams**. The amalgam used in filling teeth commonly consists of tin, silver, and mercury. *Gold alloys very readily with mercury.* Indeed, the latter readily dissolves gold, and so is used in *extracting fine particles of gold from the sands in which they occur.* Mercury is also similarly used in silver mining.

Tin, stannum, occurs in nature as **tinstone, cassiterite**, SnO_2, which is found in England, Germany, Peru, Australia, Banca, and Alaska. To obtain the tin from this ore, the latter is heated with carbon, thus:

$$SnO_2 + 2\,C = 2\,CO + Sn.$$

Tin is very malleable, melts at 232°, and boils at 1600°. On exposure to the air tin remains nearly unchanged. It is used in making **tin foil** and **tin plate**. The latter consists of sheet iron coated with tin by the process of dipping the thoroughly cleaned iron into a bath of molten tin. Frequently copper is also treated by this method. *Metallic tin is also used in*

many alloys. Thus **solder** commonly consists of 1 part tin and 1 part lead, though other proportions are also employed. The *bronzes,* already mentioned, *are alloys of tin and copper.* **Pewter** consists of 1 part lead and 3 parts tin. **Britannia metal** contains 90 per cent tin, 8 per cent antimony, and 2 per cent copper. Among the most **important salts of tin** are **stannous chloride,** $SnCl_2$, formed by dissolving tin in hot concentrated hydrochloric acid solution, and **stannic chloride,** $SnCl_4$, produced by treating tin or stannous chloride with chlorine. Stannous chloride forms white crystals that are very soluble in water, while stannic chloride is a fuming liquid which boils at 114°. With ammonium chloride stannic chloride forms a double salt, $SnCl_4 . NH_4Cl$, called **pink salt,** which is used as a mordant. Stannous chloride is a reducing agent, for it readily passes over into stannic chloride, for example:

$$\underset{\text{stannous chloride}}{SnCl_2} + \underset{\text{mercuric chloride}}{HgCl_2} = \underset{\text{stannic chloride}}{SnCl_4} + \underset{\text{mercury}}{Hg}$$

Mosaic gold is stannic sulphide, SnS_2, prepared by heating together sulphur and tin. It consists of golden yellow crystals, and is used for "bronzing" articles. *Tin is a rather costly metal.* The world produces about 110,000 tons of it annually. Most of this is mined in Banca and other neighboring East India islands.

Lead, plumbum, is soft and malleable. It melts at 327° C. and it is 11.4 times heavier than water. This metal has been used in the arts even in ancient times. The Romans used lead for water pipes, and it serves for this purpose to this day. *The chief ore of lead is the sulphide* PbS, *which crystallizes in cubes.* It is commonly known as **galenite.** As already mentioned, solder and pewter are **alloys of lead and tin.** Shot and bullets consist of lead containing about three-

tenths per cent arsenic, while in Babbitt metal from 70 to 90 per cent lead is alloyed with antimony and tin. Much lead is used also in the manufacture of lead storage batteries and in making lead salts and other compounds. The **lead storage battery** consists essentially of a lead plate and a lead plate coated with lead peroxide, PbO_2, dipping in a solution of sulphuric acid of specific gravity 1.2. As the battery acts, the current flows from the lead through the solution to the lead peroxide. Thus lead is dissolved and hydrogen is set free at the other plate, but it at once attacks the lead peroxide, forming lead oxide and water. The lead oxide formed is dissolved by the sulphuric acid present, forming water and lead sulphate. *The electromotive force of the battery is two volts.* Charging the battery consists simply in conducting a current from a dynamo through the battery from the peroxide plate through the solution to the lead plate. Thus lead is again deposited on the lead plate and lead peroxide is again formed on the opposite plate; that is to say, the chemical changes that took place when the battery was discharged are reversed. *These batteries are used for running electric automobiles, also for lighting and ignition purposes, etc.*, for they have a high electromotive force and a low internal resistance and so yield a powerful current.

FIG. 52.—A lead storage battery.

Litharge, lead oxide, PbO, is a yellow powder which is used extensively in *glazing pottery and ironware*, in putting decorations upon porcelain, and in making glass that has a high in-

dex of refraction. It is also employed in making various **lead salts**. The most common soluble lead salts are the nitrate, $Pb(NO_3)_2$, and the acetate, $Pb(C_2H_3O_2)_2 . 3\ H_2O$. The latter is commonly known as **sugar of lead**, for it has a sweetish, yet disagreeable taste. The **basic lead acetate**, $Pb(C_2H_3O_2)_2 . (PbO)_x$, is often used as a lotion for diseased parts, a 2 per cent solution of this salt being called *lead water*. **Lead arsenate**, $Pb_3(AsO_4)_2$, *is used for poisoning potato bugs and also for spraying trees and shrubs.* Like Paris green, *lead arsenate is but sparingly soluble in water.* **White lead** is a basic carbonate of lead, $Pb(OH)_2 . 2\ PbCO_3$. It is sold in large quantities ground in linseed oil. On being further diluted with the latter, *this mixture makes an excellent paint*, especially for wood that is to stand exposure to the weather. *White lead paint is not infrequently adulterated with chalk, so-called whiting, barium sulphate, or lead sulphate.* It is to be borne in mind that *all lead compounds are poisonous.* The more soluble the compounds are, the greater is the danger of being poisoned by means of them. Painters are often poisoned by the white lead they use. This gets on their hands, and it is very likely that when these touch the nose and lips, the lead compounds come into contact with the mucous membranes and are absorbed. Thus small quantities are taken into the system at a time, but these accumulate and finally cause **lead colic.**

Chromium is a hard, steel-gray, brittle metal which melts at about 1615° C. It is 6.8 times as heavy as water. It occurs in chrome iron ore, $Cr_2O_3 . FeO$, and is used in making **chrome steel**, which is steel alloyed with a small percentage of chromium. This makes a very hard steel. Chromic oxide, Cr_2O_3, is **chrome green** and is used as a pigment in paint. *It is also used in making green glass and glazes.* **Chrome yellow** is lead chromate, $PbCrO_4$. It is also used as a

pigment in paint. *The most common soluble compound of chromium is potassium bichromate*, $K_2Cr_2O_7$. It forms beautiful orange-colored crystals. *It serves as an oxidizing agent, also in chrome tanning and in dyeing fabrics.*

Tungsten is also a brittle metal. In some of its properties it resembles chromium. Tungsten is used in making **tungsten** steel, and also in producing the filaments for the incandescent **tungsten electric lamps.**

Molybdenum is closely allied to tungsten. **Ammonium molybdate**, $(NH_4)_2MoO_4$, *is used in testing for the presence of phosphates*, for when added to a nitric acid solution of the latter there forms a very characteristic yellow precipitate called *ammonium phosphomolybdate*, $(NH_4)_3PO_4 . 11\ MoO_3 . 6\ H_2O$. This precipitate is soluble in ammonia water.

Zinc is about 6.9 times as heavy as water, melts at 420° C. and boils at 918° C. It is found in nature mainly as the carbonate $ZnCO_3$, and the sulphide ZnS. The latter ore is popularly known as **blackjack**. These ores are first heated in contact with the air and are thus changed to zinc oxide.

$$\underset{\text{zinc carbonate}}{ZnCO_3} \text{ on heating } = \underset{\text{zinc oxide}}{ZnO} + \underset{\text{carbon dioxide}}{CO_2}$$

$$\underset{\text{zinc sulphide}}{ZnS} + \underset{\substack{\text{oxygen}\\\text{(from the air)}}}{3\,O} = \underset{\text{zinc oxide}}{ZnO} + \underset{\text{sulphur dioxide}}{SO_2}$$

From its oxide, zinc is then obtained by heating with carbon in earthenware retorts, thus:

$$\underset{\text{zinc oxide}}{ZnO} + \underset{\text{carbon}}{C} = \underset{\text{zinc}}{Zn} + \underset{\text{carbon monoxide}}{CO}$$

In the form of sheets *zinc is much used for gutters, roofs, ornaments, etc., on buildings.* **Galvanized iron**, so called, consists of sheet iron which has been coated with zinc by the process of dipping the thoroughly clean iron into a bath of molten

zinc. The beautiful flaky appearance of galvanized iron is due to the crystals of zinc on its surface. *The coating protects the iron so that it will not rust.* Zinc is also used in many

FIG. 53. — A sal ammoniac battery, showing the parts.

primary electrical batteries. In the common battery used for ringing doorbells, Fig. 53, there is a solution of ammonium chloride, sal ammoniac, into which dip a stick of zinc and a cylindrical plate of carbon. The latter is usually so shaped as to form also the cover of the jar. The **battery carbon** is made by grinding coke with black strap molasses or coal tar as a binder, molding, and then heating the product at first gently, and finally to redness out of contact with the air. In this battery, the current flows from the zinc through the solution to the carbon. Zinc chloride is thus formed as the zinc and hydrogen is evolved at the carbon. The hydrogen is absorbed by the carbon and finally escapes in the air. **Dry batteries** contain the same ingredients. Only here the outer casing is made of zinc, which also serves as the negative pole of the battery. The carbon generally is a cylindrical piece in the middle of the cell and is surrounded by manganese dioxide which oxidizes the hydrogen that is liberated. Plaster of Paris is used to absorb the liquid and hold the ingredients in place. But *the contents must be moist or no current can flow.* **Blue cup batteries**, Fig. 55, contain a zinc pole in dilute sulphuric acid, and a copper pole in saturated copper sulphate solution. As the current passes from zinc through the

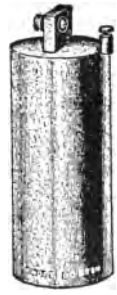


FIG. 54. — A dry battery.

solution to the copper, zinc dissolves from the zinc plate, and copper deposits on the copper plate. Other metals could be used in place of zinc, but the latter is on the whole the most economical, producing a relatively high electromotive force.

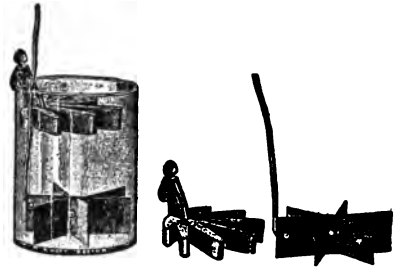


FIG. 55. — A blue cup battery and its parts.

Brass *is the most important alloy of zinc which is in common use,* but the latter metal is often present in other alloys. About two hundred and fifty thousand tons of zinc are produced in the United States each year, and the rest of the world produces about two times that amount. *The compounds of zinc are analogous to those of magnesium.* It is to be kept in mind, however, that *zinc compounds are poisonous.* **Zinc oxide**, ZnO, is used in paints, being called **zinc white.** Mixed with lard or vaseline, it forms **zinc oxide ointment.** **Zinc chloride**, $ZnCl_2$, is very soluble in water. It is used for *preserving wood,* especially railroad ties. The salt permeates the wood and is an antiseptic; thus it is death to the germs that are the cause of the decay of wood. **White vitriol** or **zinc sulphate**, $ZnSO_4 . 7\ H_2O$, is another very common zinc salt. It is readily soluble in water.

Manganese is a hard, brittle metal, which is eight times as heavy as water, and melts at 1800°. In nature it is found mainly as pyrolusite, which is **manganese dioxide**, MnO_2. *Alloys of manganese and iron are used in the steel industry.* With copper, manganese forms alloys called **manganese bronze.** These are very hard and strong. Manganese dioxide is often used in preparing chlorine, thus:

$$\underset{\text{manganese dioxide}}{MnO_2} + \underset{\text{hydrochloric acid}}{4\,HCl} = \underset{\text{water}}{2\,H_2O} + \underset{\text{manganous chloride}}{MnCl_2} + \underset{\text{chlorine}}{Cl_2}$$

The salts of manganese are quite numerous. The **manganous salts** are pink in color, and the chloride, $MnCl_2$, sulphate, $MnSO_4$, nitrate, $Mn(NO_3)_2$, and acetate, $Mn(C_2H_3O_2)_2$, are soluble in water. Manganese acts as an acidic element in *manganates* and *permanganates.* Of these salts **potassium permanganate**, $KMnO_4$, is of practical and commercial importance. This salt crystallizes in needles that are of a dark purple, lustrous hue. The salt is soluble in water and very often serves in the laboratory as an *oxidizing agent. It is also used as a disinfectant.*

Nickel is malleable, ductile, and 8.9 times as heavy as water. Its melting point is about 1485° C. **Nickel coins** consist of 25 per cent nickel and 75 per cent copper. *When alloyed with brass, nickel forms German silver.* Nickel steel is used for making **armor plates** for warships. Nickel does not tarnish readily on exposure to air, and hence it is often used to plate iron, copper, brass, etc. **Nickel plating** is carried on in the same way as gold or silver plating. The bath for nickel plating consists of a solution of **nickel ammonium sulphate**, $(NH_4)_2SO_4 . NiSO_4 . 6\ H_2O$, and, of course, a plate of solid nickel is placed in the solution opposite to the object to be plated. The electric current is then passed from the nickel plate through the solution to the object that is to receive the nickel coating. *The salts of nickel are green, and when soluble they yield green solutions.*

Cobalt is analogous to nickel in most important respects. *Its salts when dissolved in water yield dark red solutions,* and the crystals that deposit from such solutions are also red, for in general the crystals contain water of crystallization. When this water is driven off, anhydrous, *i.e.* dry, compounds

are formed which are *blue*. *Cobalt silicate is blue*, and so glass containing cobalt silicate is **blue glass**, also called **smalt glass.** Indeed, about *the only practical use which is at present made of cobalt is in the production of blue glass, blue porcelain, and blue glazes on enamel ware.* Cobalt, nickel, and iron are metals that are *attracted by a magnet*, and are consequently said to be *magnetic.*

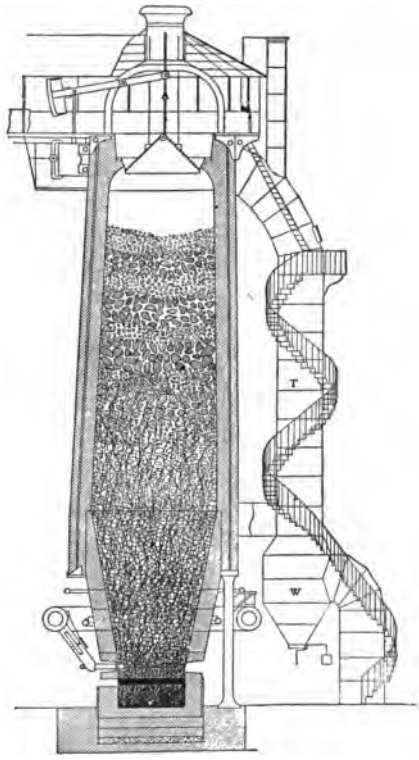


FIG. 56.—A blast furnace in which cast iron is made.

Of all the metals, **iron**, *ferrum, is the most useful.* It very rarely is found in the free state in nature except in meteorites. The most important ore of iron is **hematite**, which is an oxide having the composition Fe_2O_3. It is dark red in color and when finely ground it serves as a pigment in paint, being called **red ocher**. *The rich deposits of iron ore in the Lake Superior region consist of hematite.* **Magnetite**, or **magnetic iron ore**, Fe_3O_4, is black. **Limonite**, $2\ Fe_2O_3 . 3\ H_2O$, is a hydrous oxide of iron. It is yellow in color. It too serves as a pigment under the name of **yellow ocher**. These oxides and **siderite**, ferrous carbonate, $FeCO_3$, form the chief ores of iron.

To obtain metallic iron from these ores, they are mixed with charcoal, coke, or coal and limestone and then heated in a **blast**

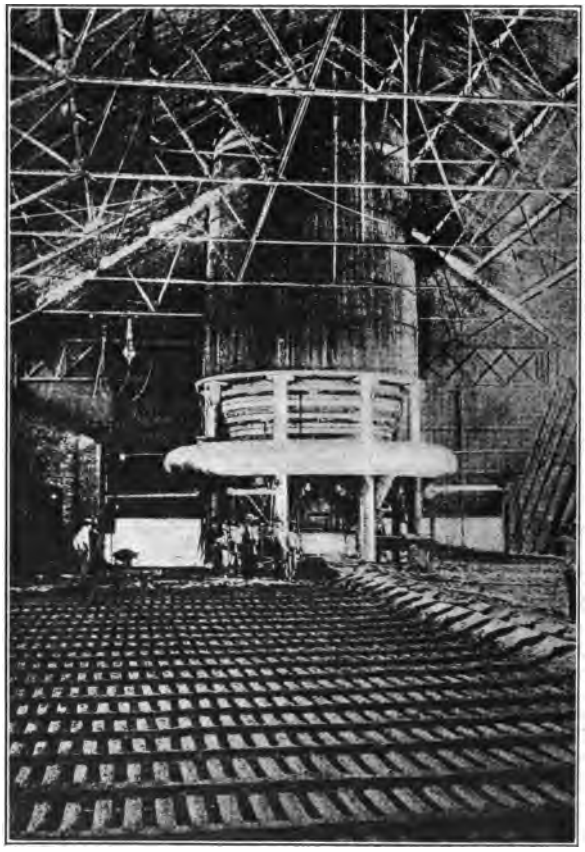


FIG. 57. — A modern blast furnace, showing the pig iron bed.

furnace. In the lower part of the furnace carbon dioxide is formed, for here air is blown in and there is consequently sufficient oxygen to form carbon dioxide. The latter gas passes upward through the hot carbon and is converted into

carbon monoxide, thus:

$$CO_2 + C = 2\ CO$$

Then the carbon monoxide reduces the iron ore as follows:

$$Fe_2O_3 + 3\ CO = 2\ Fe + 3\ CO_2$$

and out of the top of the furnace there issues carbon dioxide, mixed, however, with much carbon monoxide. This gas is now used to run gas engines that furnish the air blast for the furnace. The molten iron accumulates in the bottom of the furnace and is tapped off from time to time. The iron is run into molds of sand so as to form rough bars, or pigs as they are called. The cast iron thus obtained is known as **pig iron.** The limestone was added so that the sand and silicates might react with the lime to form a fusible calcium silicate which floats on top of the molten iron and is run off. It is the so-called **slag.** *It is now used for making Portland cement, whereas formerly it was thrown away.*

FIG. 58. — Cast iron as it appears under the microscope.

Cast iron always contains considerable amounts of carbon and other impurities. In fact *chemically pure iron is not obtainable in the market,* being practically unknown. *The different kinds of iron and steel, then, are really iron containing various other substances.* The carbon content especially determines the properties of the iron. Cast iron usually contains from 2 to 3 per cent of carbon embedded in the iron

as graphite, and from 1 to 1.5 per cent of carbon as iron carbide, *i.e.* chemically combined carbon. From 0.5 to 4 per cent of silicon, from 0.4 to 2 per cent of phosphorus, and sulphur up to 0.2 per cent are also generally present. **Wrought iron** *is the purest iron on the market.* It contains usually less than 1 per cent of foreign ingredients. The carbon content is generally less than 0.2 per cent. Wrought iron melts at 1600° C. and can be welded at from 900° to 1100°. *In welding, borax or sand is used.* Thus a slag of iron borate or iron silicate is formed, and when the iron, brought to welding heat, is pounded together, this slag flies off and the clean surfaces of the pieces of iron come into intimate contact so that the forces of cohesion can act and hold them together. *Wrought iron is made from cast iron by* **puddling**, a process which consists of adding iron oxide to the pig iron and then heating in a current of air in a so-called reverberatory furnace. Thus the carbon of the pig iron is removed because it unites with the oxygen of the iron oxide and of the air. Other impurities present, like silicon and phosphorus, are also oxidized and form a slag with some of the iron. Finally the iron is almost free from foreign matter except a few tenths of a per cent of carbon. **Malleable iron** is a cheap substitute for wrought iron. It is made by embedding the castings in iron oxide and then heating for about two days, after

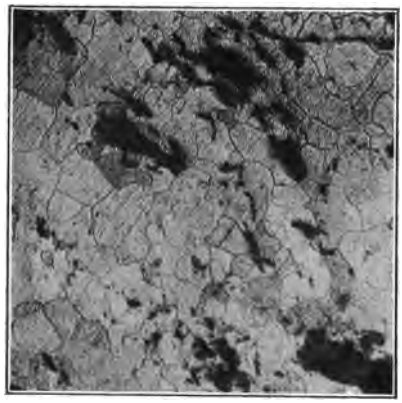


FIG. 59. — Wrought iron as it appears under the microscope.

which all is cooled off slowly. Thus some of the carbon is removed from the castings, making them less brittle than they were. It is to be borne in mind that *a large carbon content makes iron brittle. Sulphur and phosphorus are especially objectionable in iron, for they make it quite brittle.* **Steel** is practically free from sulphur and phosphorus. Silicon too is present only in small amounts. The carbon in steel varies from about 0.2 to 1.6 per cent, mild steel containing the smaller amounts of carbon. The process of **hardening** and **tempering** consists of heating the steel and then chilling it. Very sudden chilling does not give the carbon a chance to crystallize out as graphite, and *the combined carbon makes the steel very hard.* If the cooling proceeds very slowly, practically all the carbon crystallizes out, and a soft material is obtained. By suitable chilling, the proper temper, that is, the desired degree of hardness, may be obtained. *The removal of the carbon from cast iron so as to produce steel consists essentially in oxidizing the carbon.* This is accomplished by blowing air into the molten mass, as in the **Bessemer process**, in which a so-called Bessemer converter is used, or by heating cast iron with iron oxide on the hearth of a furnace, the so-called **open hearth process**. When the cast iron contains much phosphorus, the hearth of the furnace is made of lime

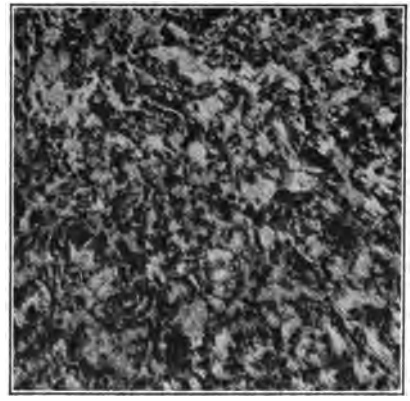


FIG. 60. — Steel as it appears under the microscope.

obtained by calcining dolomite. This absorbs the phosphorus, phosphates of calcium and magnesium being formed which are sold as fertilizer. *Nearly thirty million tons of iron are produced yearly in the United States.*

FIG. 61. — An open hearth furnace.

Iron forms two series of salts, the ferrous and the ferric. Thus there are **ferrous chloride**, $FeCl_2$, and **ferric chloride**, $FeCl_3$, **ferrous oxide**, FeO, and **ferric oxide**, Fe_2O_3, etc. In the **rusting of iron** hydrated ferric oxide is formed. The black oxide, also known popularly as **hammer black**, which results when iron is heated in the air, is **ferrous ferric oxide**, $FeO \cdot Fe_2O_3$ or Fe_3O_4. It is magnetic. Iron nearly always acts as a basic element, and in this capacity it forms salts with practically all of the various acids. Only a few of the most important salts of iron will be mentioned here.

Ferrous sulphate, also known as **green vitriol** or **copperas**, is $FeSO_4 \cdot 7\ H_2O$. It is formed by the oxidation of **pyrite**,

FeS_2, which occurs in nature and is called **fool's gold**, for its crystals are of a lustrous, golden yellow appearance. Ferrous sulphate dissolves readily in water. *It is used in making ink.* The latter contains tannin and ferrous sulphate. The **tannin** is an essential ingredient in extract of nutgalls, which is added to the ferrous sulphate solution. Ferrous sulphate is also used as a disinfectant, as a mordant in dyeing fabrics, and as a reducing agent. This salt is cheap and readily obtainable from dealers everywhere.

Ferric chloride, $FeCl_3 . 6\ H_2O$, forms a dark brown crystalline mass, not unlike maple sugar in appearance. It dissolves copiously in water, also in alcohol and ether. It is used in medicine. The so-called **styptic cotton** which is used *to stop the bleeding of wounds* consists of absorbent cotton treated with a solution of ferric chloride. This material not only allays bleeding by forming a clot with the blood, but it also acts as an antiseptic, thus protecting the wound from germs.

Blue print paper consists of paper that has been coated with a solution of *ferric ammonium citrate* plus *potassium ferric cyanide.* The latter is also called **red prussiate of potash**, $K_3Fe(CN)_6$. This paper is thus coated and dried in the dark. On exposure to light, the ferric citrate is in part reduced to ferrous citrate, and the latter, like other soluble ferrous salts, reacts with the potassium ferric cyanide, forming a blue precipitate, $Fe_3''\,[Fe'''\,(CN)_6]_2$, Turnbull's blue. Where the paper has been protected, no precipitate forms; and so when after exposure to light in the printing process the paper is washed with water, the parts on which the light has acted appear blue, whereas the parts that have been protected appear white, for the original coating on the blue print paper is soluble in water. It is to be remembered that if blue print paper, not exposed to light, is washed with water in the dark, white paper is obtained. *The blue printing is done in frames*

similar to those used for printing photographs from negatives. The latter will also serve for making blue prints. But blue prints are commonly made by using tracing cloth or paper that is not too thick on which the writing or drawing that is to be reproduced in blue print form has been traced. The thicker the paper, the longer must be the exposure to the light to get the desired results. *Ammonia, caustic soda, and caustic potash solutions decompose Turnbull's blue, and so may be used to write white characters on blue prints.* The latter are stable in the light, also toward acids, but alkalies, as just mentioned, destroy the blue color.

It is to be borne in mind that *in small amounts iron is present almost everywhere.* In all soils it occurs. To the brown sand, earth, and the brownish and yellowish clays it gives their characteristic colors. Our sandstones and other rocks all contain compounds of iron. Indeed, *the grains of sand in the sandstones are commonly cemented together with oxides of iron. Iron compounds are present in the tissues of all plants and animals.* The green leaf contains *chlorophyl,* for the formation of which iron is essential, and the blood of animals contains *hæmoglobin,* in which iron plays an important rôle. In fact, *without iron plants and animals cannot live.* Nevertheless, it must be remembered that the quantity of iron present in living beings, exceedingly important though it is, is after all small; so, for instance, the human body contains only about 0.004 per cent of iron.

QUESTIONS

1. Where is aluminum found in nature, and in what form?
2. How is metallic aluminum made?
3. Mention the most important characteristics of metallic aluminum. What is it used for?
4. What is thermite? Describe its use.

5. What is alum, and what is it used for?

6. How is porcelain made?

7. What is the difference between porcelain and ordinary crockery ware?

8. How is ultramarine made? What are its uses?

9. How much alumina can be made from 45 grams of potassium alum?

10. What are the noble metals? Describe each.

11. For what purposes is gold used? Why?

12. What is the most important compound of silver? What are its uses?

13. What is sterling silver? What is it used for?

14. Discuss the use of silver compounds in photography.

15. What is a negative?

16. Where does copper occur in nature, and what are its chief characteristics and uses?

17. Mention three important alloys of copper.

18. Describe the most common compound of copper.

19. What is Bordeaux mixture?

20. Of what use is mercury?

21. Mention two important compounds of mercury and state what they are used for.

22. How is tin obtained from its ores?

23. What is pewter? Britannia metal? Solder?

24. What is pink salt and what is it used for?

25. What is mosaic gold? State its use.

26. Mention the properties and uses of lead.

27. What does a lead storage battery consist of?

28. What is lead arsenate? Sugar of lead? White lead? What is each used for?

29. What use is made of potassium bichromate?

30. What is ammonium molybdate used for?

31. How does zinc occur in nature? How is it obtained from these ores?

32. What is galvanized iron?

33. Describe an ordinary battery such as is used for ringing a doorbell.

34. State the uses of zinc oxide and zinc chloride.

35. What is manganese dioxide used for? Write the equation expressing the chemical changes.

36. Of what use is potassium permanganate?

37. State the characteristics and uses of nickel.

38. How proceed to nickel plate an iron spoon?

39. What use is made of cobalt?

40. Name the chief iron ores and state briefly how iron is obtained from its ores.

41. What is the chief difference between cast iron and wrought iron?

42. How is steel made, and how is it tempered?

43. Why are sulphur and phosphorus objectionable in iron?

44. What is malleable iron? How is it made?

45. Explain how fertilizers and also Portland cement are made in connection with the production of iron and steel.

46. How many series of salts does iron form? Mention two compounds of each series.

47. What is ordinary ink?

48. Describe how to make a blue print.

49. Discuss the occurrence of iron in plants and animals.

CHAPTER XII

PAINTS, OILS, AND VARNISHES

THE use of paints, oils, varnishes, and lacquers dates back to ancient times, for the great advantage of putting a protective coating upon articles of wood, metal, etc., has been appreciated for many centuries. That such coatings frequently enhance the beauty of the objects to which they are applied was also duly recognized.

It has already been stated that *oils may be divided into two great classes:* (1) *the mineral oils, and* (2) *the oils of organic, that is to say, of plant or animal, origin.* But oils may also be classified according to their use. *When an oil is to be used in paint it must be an oil that will dry and form a hard coating.* Oils like *lard oil, olive oil, cottonseed oil, goose oil, neat's-foot oil, and the various mineral oils* obtained from petroleum *will not dry.* They remain smeary when they are spread upon wood, glass, metal, etc., and they are consequently called **non-drying oils.** On the other hand, **linseed oil** and **poppy-seed oil,** for example, will, when similarly spread out upon objects, gradually dry to a hard, firm, resistant coating. Such oils are termed **drying oils.** While mineral oils are obtained from the distillation of crude petroleum, animal oils are obtained by heating animal fats and then subjecting them to pressure, and vegetable oils are prepared by grinding seeds and expressing the oil from them either at room temperature or at higher temperatures. Sometimes solvents like carbon bisulphide and gasoline are used in extracting animal or vegetable oils from fatty animal or plant

matter. Thus the fats pass into the solvent, and the clear solution containing the fat may be filtered from the solid residue that remains. From the solution the oil or fat is obtained by distilling off the solvent, which can be used over and over again.

FIG. 62. — A press used in removing oil from seeds.

Linseed oil is the most important of all oils for use in paints. It is obtained from flaxseed, one bushel of the latter yielding about 2.3 gallons of oil. *The drying qualities of linseed oil depend upon the fact that it has the power to absorb oxygen from the air.* This oxidized linseed oil, also called linoxin, forms a tough, hard, resistant film and is quite unlike the original oil. *By boiling linseed oil with lead oxide, oxides of manganese, or linoleates of these metals, the oil dries more rapidly,* and the substances used to bring about this effect are called **driers.** They are, in general, oxidizing agents whose purpose is to increase the rapidity of the oxidation of linseed oil to linoxin. **Linseed oil** itself is essentially a glycerine salt of *linoleic acid,* of the formula $(C_{18}H_{31}O)_3 . C_3H_5O_3$. There are also present in the oil, to the extent of about 20 per cent, olein, palmitin, etc. American linseed oil has a specific gravity of 0.9336 at 15° when raw, whereas the boiled has a specific gravity of 0.938. It boils at 130° C. *The raw oil obtained by pressing ground flaxseed in the cold is very light-colored, whereas oil that is pressed from the hot seeds, and also boiled oil, is dark brown in color and has a greenish tinge.*

Upwards of fifteen million bushels of flaxseed are raised annually in the United States, which would represent about *thirty-four million gallons of linseed oil.* Being somewhat expensive, linseed oil is often shamefully adulterated with cheaper oils, among which may be mentioned fish oil, petroleum oils, rape, cotton, hemp, and corn oil. The presence of these adulterants, of course, lowers the quality of the paint made by using the admixture. Rosin is also frequently used to adulterate linseed oil.

When linseed oil is thoroughly mixed with white lead, an excellent paint is obtained. This is usually slightly yellowish in tint, due to the color of the oil, but by adding a very small amount of blue — Prussian blue, for example — this yellow shade is dispelled and a pure white paint is obtained, which

FIG. 63. — A barrel of white lead.

dries rather slowly, but is exceedingly durable, especially for outside work that is to be exposed to the weather. The paint can be made to dry more rapidly by adding small amounts of so-called *driers*, sometimes also termed *Japan driers*. As stated before, they consist of substances that absorb oxygen from the air and then give it off to the oil, at

least in part, thus hastening the oxidation, that is, the drying, of the oil. **Turpentine** is a substance of this kind, and so its addition to paint hastens the process of drying. Driers are commonly made by heating lead and manganese oxides (about four pounds of the mixture) with a gallon of linseed oil to about 500° F. and stirring till a uniform mass results. This is then dissolved in turpentine before it is cold. By adding some of the solution thus obtained to paint, the latter dries rapidly. *Too much drier must not be used, or the paint after hardening will keep on oxidizing, and thus the oil will be oxidized too much and spoiled,* or, as the painters say, "burnt." Any drier really tends to make the paint less durable. There is much linseed oil sold as boiled oil which is *not* boiled at all, being *raw oil to which drier has been added;* it is popularly called "bunghole boiled oil." Often zinc oxide is used in paints together with white lead, even to the extent of half and half. *Paints may be colored by adding suitable pigments.* These should be finely ground and then thoroughly mixed with the white lead and oil. Stirring is really not as good as grinding all together. The **pigments** remain suspended in the form of minute particles in the paint and also in the film as it dries after it has been applied. As already stated, **yellow ocher** and **red ocher** are oxides of iron. They are cheap, permanent, and excellent pigments. **Chrome yellow** gives a beautiful canary-yellow shade, but it is more costly. **Vermilion** is suphide of mercury, cinnabar; and **chrome green,** which is very brilliant in hue, is chromic oxide. **Prussian blue** is produced by adding ferric chloride to a solution of potassium ferrocyanide, $K_4Fe(CN)_6$. It is often used as a pigment. **Ultramarine blue** too is used, but it is not as permanent. *It is excellent for interior work in kalsomines, etc.* Shades of gray result by adding small quantities of lampblack to white paint. *It is always best*

to apply two or more thin coats of paint well brushed out, rather than to attempt to cover the object with but one thick coat. This is obvious, for a thick coat will not dry readily, and will tend to peel off. As adulterants of white lead, chalk, barytes, lead sulphate, kaolin, and a mixture of zinc sulphide and barytes, which is termed **lithophone**, are frequently used. None of these should be present in high grade paint.

Varnish *is made by carefully melting rosin and then stirring in linseed oil and heating the mixture together.* This is finally

FIG. 64. — Gathering turpentine.

thinned to the proper consistency with turpentine. **Turpentine**, $C_{10}H_{16}$, is obtained by distilling the sap collected from pine trees. After the turpentine has been distilled off,

the **rosin** remains behind in the retort. Many paints contain varnish, which gives a gloss to the film that forms on drying. Sometimes benzine or gasoline are used to thin varnishes or paints instead of turpentine, but this is bad practice, for the film that remains after the benzine has evaporated is not durable. The turpentine never wholly evaporates; it always leaves a film of its own which together with the varnish or paint film makes for greater durability. If, instead of rosin, **copal**, a fossil rosin found in Africa, is used in making varnish, a far better article is obtained, but it is also much more costly. **White enamel paints** are superior varnish to which white lead or zinc oxide or both have been added as a pigment. Other **colored enamel paints** result by introducing suitable pigments.

Shellac varnish consists of a solution of *shellac*, a plant resin, in alcohol. Either wood alcohol or grain alcohol may be used. *White shellac* is bleached shellac, whereas the *brown or golden shellac* has the natural color of the resin. Shellac varnish is cheaper, but not as durable, as varnish made from oil and rosin.

FIG. 65. — Trinidad asphalt lake near the edge.

Black varnishes used for coating iron are commonly prepared by dissolving coal-tar pitch or asphaltum in turpentine or benzine or a mixture of the two. **Asphalt**, also called **asphaltum**, is a natural pitch. Large quantities of it are found in Trinidad. It is also used for making excellent street pavements, by mixing it with sand, dust, and finely crushed stone and then applying this on a foundation of concrete, rolling the hot mixture securely into place by means of heavy, hot steam rollers.

FIG. 66. — An asphalt street in New York.

Kalsomines consist of a thin solution of glue to which slaked lime or chalk or both have been added. These mixtures may be tinted by means of pigments just as in the case of paints. Many kalsomines contain plaster of Paris, which forms a hard coating as it dries. *Kalsomines, being soluble in water, can be used only for indoor work,* preferably on plastered walls.

QUESTIONS

1. For what purposes are paints and varnishes used?
2. How may oils be classified?
3. What is a drying oil? Name one and state why it dries.
4. How are plant oils obtained? Name several plant oils.
5. How are animal oils prepared? Name several. Are these suitable for use in paint? Why?
6. What is a so-called drier? How are driers used?
7. What are some of the adulterants used in linseed oil?
8. Why is turpentine better than benzine in paints and varnishes?
9. Describe how to make a good paint for outside woodwork.
10. Mention five pigments that are used in paints.
11. How is varnish made?
12. What is shellac varnish?
13. What is black varnish and for what purpose is it used?
14. What is asphalt? For what is it used?
15. What is kalsomine?

CHAPTER XIII

LEATHER, SILK, WOOL, COTTON, AND RUBBER

WEARING apparel is made of materials derived from animals or plants. Shoes are made of leather, which is produced

FIG. 67. — Silkworm, cocoon, egg-mass, and strands of silk.

from the skin of animals. Furs, too, are skins of animals that have been subjected to certain treatment. **Silk** is made from the cocoon of the silkworm. The silk threads are spun and then woven into cloth. **Wool** is the hair of the sheep. It is cleansed, freed from fat, spun, and then woven. *Wool and silk are substances containing carbon, hydrogen, oxygen, sulphur, and*

FIG. 68. — A sheep.

nitrogen. They are very like horn, hoofs, and finger nails in chemical composition, and *when burnt all of these give a peculiar odor, namely, the odor of burnt hair.* This odor is due to the nitrogenous decomposition products that are formed. Thus it is quite easy to distinguish between wool and silk on the one hand, and cotton on the other, for the latter does *not* evolve the odor of burnt hair when ignited. **Cotton,** as has already been pointed out, is practically pure cellulose $(C_6H_{10}O_5)_x$, and so contains no nitrogen. *Moreover, wool*

and silk are readily soluble in caustic alkalies, whereas cotton is not. Cotton, too, is freed from mechanical impurities, and then spun and woven. *Because of the great chemical differences between cotton and wool, they react somewhat differently toward dyestuffs;* and if a fabric containing both cotton and woolen threads is dyed in the same liquid, their threads will appear of different shades. Frequently cloth is made by using cotton warp and woolen woof, and the rather rough bulky nature of the woolen threads after the process of milling

FIG. 69. — A fleece of wool.

hides the cotton threads. Such cloth is often quite serviceable, though it is much cheaper than cloth that is all wool, for cotton is far cheaper than wool. Old woolen clothing is frequently taken apart and used over again to make new cloth. In this process the threads of wool are necessarily much shortened and the cloth woven therefrom is much less durable than new goods. By using some cotton in the warp of such **made over woolens**, a better wearing cloth is obtained. When used over again and again, however, the threads finally become so short as to be worthless except as a fertilizer.

Moreover, when used for a second time, there is cotton mixed with the woolen threads, and such material woven on cotton warp is commonly known as **shoddy**. *Silk, woolen, cotton, and linen fabrics are now dyed almost exclusively by means of*

FIG. 70. — An old-fashioned loom of colonial days.

the aniline dyestuffs, also called coal-tar colors. As a rule the colors are more readily fixed upon silk and woolen fibers, and less readily upon the cotton and linen fibers. **Linen** is made of flax. Like cotton, it is essentially cellulose. **Mordants** *often have to be used to fix dyestuffs on cotton and linen, and sometimes also upon wool and silk.* **Dyestuffs** are in general either slightly acid or basic in character. Mordants, too, are

either weakly acidic or basic substances. An illustration will serve to show **the use of a mordant**. Suppose the mordant is tannic acid; by first dipping the fabric into a solution of this acid, the fibers take up a certain amount of the substance; and when they are now immersed in a dye of basic properties, the latter is fixed upon the fiber.

In making **leather,** *the object is to produce a material that is strong, pliable, insoluble in water, and not subject to putrefaction.* To this end the hides or skins are first thoroughly softened by soaking them in water. They are then immersed in a bath of slaked lime (milk of lime), which loosens the hair so that it can afterward be readily removed mechanically. The process of removing the hair is called **depilation.** It is also sometimes accomplished by simply hanging the moist skins up so that slight putrefaction ensues, whereupon the hair becomes loose and can readily be removed. If the lime process of depilation has been employed, the lime in the skin must be removed before proceeding farther. This is done by treating in a bath of dilute sulphuric acid and then washing with water. This treatment greatly increases the bulk of the skin, and thus it is in excellent condition for the next step. This consists of immersion in *extract of hemlock or oak bark. This extract contains* **tannin** *which unites chemically with the fibers of the hide and produces the compound which is called leather.* Leather does not dissolve in water and will not putrefy. It requires about seventy days to convert the hide into leather by immersing the hide in extract of tanbark as described. Often hides are put into vats with alternate layers of ground tanbark, and water is then added to cover all. Thus the water extracts tannin from the bark, and leather is formed by the union of the hide with the tannin. This process requires a much longer time. Two or three months, or even a year, may be required to secure the end in

view. Moreover, the bark becomes exhausted and the hides have to be placed in another vat with new bark from time to time. Nevertheless, in this way an excellent product is obtained. *A hide gains about 35 per cent in weight during the process of tanning.* Other barks besides oak or hemlock are also used. So, for example, *sumach* is employed in making **morocco leather** from goatskins or **imitation morocco** from

FIG. 71. — A scene in a tannery. Scrubbing the hides.

sheepskins. The tanning in this case proceeds rapidly; usually it lasts only a day or two. The hair of the goatskins is removed by the lime process, but the excess of lime is then got rid of by using hen or pigeon manure. *This keeps the leather soft.* In other cases **soft leathers** are secured by removing the lime by means of sour bran prepared by treat-

ing bran with water and sour dough. These methods of removing the lime are called "**bating**," for by employing infusions of the materials named the lime is removed and a large swelling of the hide is "abated." *The* **bate liquor** *contains organic acids which form soluble salts with the lime, and these then can be washed out.* Morocco leather is dyed with aniline dyes. **Russia leather** is produced in a similar manner, only the tanning is done with *willow or tamarack bark.* **Kid leather,** such as is used for gloves, is made of goatskins or sheepskins. These are unhaired with lime, and then treated with sour bran to remove the excess of the lime. The actual tanning consists of treating the skins with a mixture of common salt, alum, flour, and egg yolk in water. This is the so-called **tawing process.** The alum unites with the fibers of the skin. It is moreover antiseptic and so guards against putrefaction. The result obtained is a white, very soft, pliable leather.

When the process of tanning is carried so far that the product no longer yields gelatine on being boiled in water, we have **sole leather.** The ordinary soft leathers yield some gelatine on boiling in water; but kid leather and **chamois skin** or **buckskin** do not. The latter are made from skins of deer, goats, or sheep in much the same way as morocco leather, only they are finally treated with animal oils to make them very soft and pliable. The excess of oil is removed by means of alkalies. **Split leather** is made by splitting thick hides, such as cowhide, by means of machinery. Three or four or more thin layers are thus made from one skin. These are then tanned separately and the grain is pressed upon them by machines. This *split leather is cheap, but not as durable as that from natural thin skins.*

Much leather is tanned in America by the **chrome-tanning process.** The hides are unhaired and cleaned in the usual

manner. They are then soaked in a dilute solution of hydrochloric acid and potassium bichromate, after which they are placed in a vat containing a reducing solution like sodium sulphite, for example. Thus *hydrated chromic oxide is precipitated, which unites with the fibers of the hide, much the same way that alumina does in making leather for kid gloves.* The chrome tanning is completed in the course of a few hours. The leather is finally washed, dyed, and oiled as leather made with tannin. **Box calf leather** is chrome-tanned. Its cut edges are dark green in color, due to the presence of chromic oxide.

Skins that are to be used for **furs** *are tanned by the* **alum process**. They are cleansed, dried, oiled, treated with sour bran and water, and finally tawed with a solution of common salt and alum. All *alum-tanned leather when thoroughly wet gets hard when it dries,* for water dissolves out the alum, and the leather then dries hard like rawhide.

Rubber is produced from the milky juice or so-called **latex** of the india-rubber trees, which grow in tropical climes. The best rubber comes from Brazil, and is called **Para rubber.** The milky juice is obtained from the trees by making incisions in the bark and collecting the exuding fluid in suitable vessels. About six ounces of latex are collected from a tree in the course of three days, and this yields about 1.8 ounces of rubber. *A tree yields about ten pounds of rubber per year.* The world produces about seventy-five thousand tons of rubber annually, of which supply about one-half comes from Central and South America, the tropics of Africa and Asia furnishing the other half in approximately equal shares. The island of Ceylon alone produces about 3000 tons per year. While the bulk of the rubber is still obtained from trees that grow wild in tropical forests, a fair share is already supplied by cultivated trees of **rubber plantations.**

The **method of preparing rubber** from the latex is still rather primitive. After the juice has been collected, it is spread upon a piece of wood, about four feet long, shaped like a paddle. Usually the paddle is simply dipped into the liquid. It is then dried by holding the paddle over a fire so that the smoke comes into contact with the drying latex. Frequently the shells of oily tropical nuts are used as fuel, and the creosote and oily vapors contained in the smoke of the fire permeate the material on the paddle and help to preserve it. When the first layer is dry, the paddle is again wet with the juice. This in turn is dried on over the fire, and so on, till many layers have been formed and a thick covering has been obtained. This is then slitted and stripped from the paddle. It is called **caoutchouc**, and is placed on the market. *This material is crude rubber. It is essentially a hydrocarbon corresponding to the composition* $(C_5H_8)_x$, and is probaby a condensation product of the hydrocarbon isoprene, C_5H_8. Turpentine, $C_{10}H_{16}$, is a related substance.

FIG. 72. — One method of tapping rubber trees.

Pure Para rubber consists of about 94.6 per cent caoutchouc, 0.14 per cent ash, 0.85 per cent water, 2.66 per cent resin, and 1.75 per cent protein substances. It is very elastic and not readily attacked by ordinary chemical agents. *It does very gradually undergo oxidation on exposure to the air, and the oxygen thus taken up unites with the caoutchouc to form a hard, brittle compound.* Thus it is that *all rubber gradually deteriorates, whether in use or not.* The rubber goods in the market are practically all made of so-called **vulcanized rubber.** The **process of vulcanizing rubber** was discovered by Goodyear in 1843. It consists essentially of adding about 10 per cent of sulphur to the caoutchouc, then molding the desired article from this mixture, and heating out of contact with the air to 140° to 150° C. A portion of the sulphur unites chemically with the caoutchouc; the remainder, however, is nevertheless required to produce the desired product. Such *vulcanized rubber is more elastic, less porous, and not at all sticky* as compared with caoutchouc. It is also less soluble in solvents like carbon bisulphide, turpentine, benzine, and naphtha, which are the **common solvents for caoutchouc.** Furthermore, vulcanized rubber does not lose its elasticity when subjected to cold. Sometimes rubber is vulcanized without heating, by simply treating it with chloride of sulphur, which is a liquid at room temperatures. This is, however, not as satisfactory for many purposes. **White rubber** articles are made by incorporating chalk or kaolin mechanically with the rubber. **Red rubber** is similarly colored by using antimony sulphide, **black rubber** by means of lampblack, **blue rubber** by using ultramarine blue, and so on. **Rubber coats, boots, shoes,** etc., are made by spreading upon the cloth fabrics a plastic layer of caoutchouc properly mixed with sulphur. The vulcanizing is then accomplished by heating to 140° to 150° C. Treating with a solution of

chloride of sulphur in carbon disulphide is also in vogue, but it does not yield as good results.

Hard rubber, also called **vulcanite** or **ebonite**, is obtained when caoutchouc is mixed with about thirty to fifty per cent of sulphur and then heated to temperatures that are from thirty to forty degrees higher than those employed in making soft vulcanized rubber. Hard rubber is used in making combs, buttons, knife handles, bowls, trays, and many other useful articles.

Rubber is a non-conductor of electricity and hence is much used as an insulating material for electrical wires. Hard rubber, too, is used as an insulator on electrical machinery and appliances. **Gutta-percha**, which comes from the gum of the percha tree, is chemically very similar to rubber. It is, however, not as elastic. *Being a good insulator, it is much used for covering wires and cables.* It also serves like hard rubber for the manufacture of various articles.

The demand for rubber has increased enormously of recent years on account of the use of rubber tires on automobiles and other vehicles. Attempts have consequently been made to make rubber artificially. While these have been successful, the **synthetic rubber** thus produced is nevertheless still so costly that it is not yet made on a commercial scale. **Rubber substitutes** are, however, on the market. These consist essentially of products obtained by heating certain oils, notably corn oil, to higher temperatures, where partial decomposition ensues and a somewhat elastic mass is obtained, which is worked up with sulphur. Such rubber substitutes are not at all the equal of real rubber, and they are in general only employed when mixed with some of the genuine article.

QUESTIONS

1. What are the different sources from which the material for making wearing apparel is obtained?
2. How distinguish between wool and cotton?
3. What is shoddy?
4. What is linen?
5. How are fabrics dyed?
6. What is a mordant? Describe its use.
7. Describe the process of making leather by use of oak tanbark?
8. What is sole leather?
9. What is split leather?
10. How is leather for white kid gloves made?
11. How is "buckskin" leather made?
12. How are skins that are to be used for furs treated?
13. Describe the process of chrome tanning.
14. How is rubber prepared from the rubber trees?
15. What is meant by vulcanizing rubber? How does this process improve the material?
16. How is hard rubber obtained? What is it used for?
17. What is gutta-percha, and what are its uses?

CHAPTER XIV

THE SOIL

LARGE areas of the earth's surface are covered with a loose layer of sand and clay mixed with the remains of plants that have decayed. *This layer is called* **soil**. Its depth varies greatly, being usually only from six to twelve inches, though soils may at times be several feet deep, as the experience that has been gathered incidentally in digging wells has shown. *Living organisms of various kinds and variable amounts of water and air are also present in soils.* The soil, which is the surface layer, rests upon the subsoil, which differs from the upper layer in that it contains less organic matter and is generally lighter in color. The subsoil gradually grades off into the solid rock below, as is illustrated by Fig. 73.

FIG. 73. — Soil formation showing the transition of rock to soil.

Sand and clay, which form by far the larger portion of soils, are mineral matter consisting of particles of silica, silicates, and frequently also carbonates, that have been broken off from large masses of solid rock. This breaking down or disintegration of solid rock into small fragments is very important in the process of **soil formation**. It proceeds today just as it has been going on in the past. *The chief agencies which are active in the process of making soil are water, wind, animals, and plants.* The effect of each of these will now be considered.

The work of water and the air in disintegrating rocks is commonly termed **weathering of rock.** It proceeds constantly. The action of water and air is in part purely physical, that is, mechanical, and again it is partly chemical in its nature. All rocks are more or less porous and consequently allow water to soak into them. As this freezes, it expands and bursts the rocks, which then gradually crumble. Again, the water dissolves and carries away some of the material of which rocks are composed. This effect is chemical in character, and it further weakens the structure of the rock and also reduces the size of the particles.

The waters of shallow streams roll the gravel stones along their beds. As these pebbles are carried along, they strike against one another; their corners are knocked off; they are ground smooth by the continual rubbing; and at the same time they are also reduced in size by the solvent action of the water. Thus the particles become smaller and smaller till they are nothing but sand. These finer particles then are generally the less soluble portions of the disintegrated rocks, and as they are carried forward by the stream they finally arrive in still water, where they sink to the bottom. When the stream eventually changes its course, they appear as real sand. In a similar way the waves on the shores of

lakes and seas are continually reducing rocks to fine particles.

In the form of ice, water is a very important mechanical agent in the disintegration of rocks. So the cakes of floating ice as they are carried down by rapid streams in spring are commonly covered with mud, rocks, and not infrequently portions of trees and other detritus that has been torn away from the banks. The ice on the shores of lakes incloses bowlders and shoves them upon the beach during the spring. Geologists have gathered evidence which shows that centuries ago the North American continent, about as far south as the Ohio River, was covered with ice to a great depth; see map, Fig. 74. These immense **glaciers** moved southward like great rivers, only extremely slowly. Their movement was doubtless entirely similar to that of the glaciers that exist at the present time in Switzerland, Greenland, and in the Rocky Mountains. This large field of glacial ice carried everything before it. It gathered up great rocks, which froze into the ice, and thus this heavy rock-shod surface slowly moved over the underlying rocks, crushing and grinding them to fine powder, which is now the sand and clay of the regions that have thus been visited by the glaciers. The latter have thus been invaluable in preparing the land for agricultural purposes. All over the North where the glaciers have passed may be found huge **bowlders** that escaped crushing and were left behind when the ice melted. These are of very hard material and so are commonly called "hard heads." They are round and smooth because of the grinding and rolling action to which they were subjected. It is clear, then, that *the soils of these glaciated regions have really been transported to their present localities by the glaciers.* In the South, on the other hand, *where the glaciers have not been, the soils have not been transported.* Here they have been formed from the

underlying rock upon which they rest to-day. Such soils that have been formed by the weathering of the particular rock upon which they rest are called **sedentary soils**, whereas

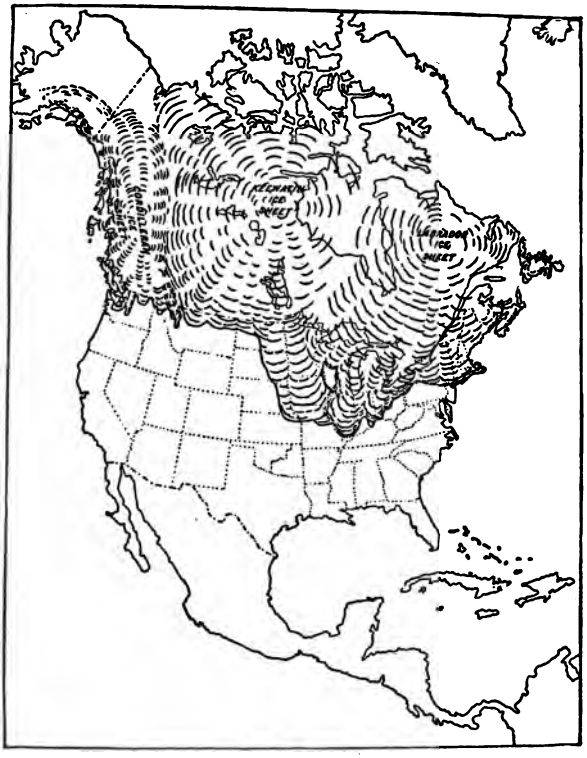


FIG. 74. — Map of North America, showing the southern limit of the glaciers.

soils that have been carried from other localities to the place where they are found are termed **transported soils**. Thus the soils made by glaciers belong to the latter class. But transported soils are also found in regions that have not been glaciated, for wind and water are active everywhere in transporting soil material. Thus the rich soil in the lower

regions of river valleys consists largely of material which has been brought down by the river from the higher portions of the valley; and since this material has in many cases been brought from various rock formations, the resulting soil generally possesses a greater fertility than if it had been formed by the weathering of any one kind of rock. The rich **bottom lands** of many of our rivers are excellent examples of this. *A good overflow from a river is often as beneficial as a covering of manure.* It should, however, be remembered that *soil once formed is liable to be washed away.* Soils composed of very fine material are thus carried off most readily, because the rain water does not soak into them rapidly, but accumulates on the surface and runs off in little streams which carry the fine topsoil with them. This is a decided loss for the farmer, for it is the best part of the soil that is thus washed away.

In the formation of soils, **the work of the wind** consists largely in carrying and distributing particles of the soil; but it nevertheless also takes part in the process of actually disintegrating rocks and grinding coarser grains to finer ones. So when particles of sand are blown against solid rock, the latter is gradually worn away and becomes soil. The action of the wind in blowing particles of sand against rock surfaces is comparable with the effect produced by the water of a stream as it causes sand and pebbles to grind on the bottom of its bed. **Sand dunes**, like those along the shores of Lake Michigan, for example, are good illustrations of wind-borne materials. While these particular dunes are worthless for farming, there are nevertheless large areas of **wind-formed soils**, as, for example, the **loess** just west of the Mississippi River, which are quite fertile and deep. *The air also acts chemically in the process of soil formation.* Many rocks and rock particles that contain ferrous silicates are oxidized to

ferric compounds by the oxygen of the air, and thus farther disintegration results. This action is especially facilitated by water, which always contains oxygen in solution. *Carbon dioxide, too, is always present in water that has been in contact*

FIG. 75. — Sand dunes on the shore of Lake Michigan.

with the air, and this dissolves many rocks, especially those that consist of the carbonates of calcium, magnesium, iron, etc.

Animals are also important agents in soil formation. So, for example, rabbits, moles, prairie dogs, and other burrowing animals dig into the earth and throw out raw subsoil and pieces of rock, all of which help to make new soil through the further action of alternate freezing and thawing. Even the action of those humble creatures, the **earthworms**, is of material consequence in the work of forming soils. They bring a portion of the subsoil to the surface, draw dead leaves and other vegetable material into their burrows, and pass large quantities of the soil through their bodies, depositing it on the surface at a rate estimated as ten tons per

acre annually. **Ants** are also active soil formers, and in some warm climates, as in Africa, they perform much the same work as the earthworm.

Growing plants send their roots into rocks and soils, thus loosening them up and rendering them porous, a condition that allows air and water to enter more readily. Roots sometimes penetrate the soil to great depth, and as they decay after the death of the plant, they leave in the soil little channels through which water gains access to the lower strata. This water always contains carbon dioxide from the air; and as already stated, it dissolves portions of the soil and thus makes them available as food for the rootlets of new plants. The roots and root hairs of living plants also give off carbon dioxide, which is absorbed by the soil water and then exercises similar solvent action upon the soil. Indeed, this *carbonated water in the soil is probably the chief agent that dissolves rocks and soil materials*, thus enabling plants to obtain the necessary elements for their growth. It must not be forgotten that all of these **effects of plants in rock disintegration and soil formation** *are slight in any one year; they nevertheless amount to a great deal when carried on through hundreds of years.* A small crack in a rock enables the tiny plants to work their way into it. As they grow, they exert great pressure, which is frequently sufficient to split the rock asunder. It is not uncommon to see bowlders weighing several tons split in two and a tree growing up between the pieces.

Peat bogs are formed in swampy places or shallow ponds where plants grow year after year, dying at the end of the season, and falling to the ground or into the water. As this process is repeated for many years, there finally results a large accumulation of vegetable material. This rots more or less and so comes to be a type of soil. If the rotting takes place under water, it proceeds very slowly and results in the

formation of peat, which is fairly solid and generally shows some of the original shapes of the plants. As stated in Chapter IX, *this material is in the first stage toward the formation of coal.* In the low and undrained marshes of this country there are countless tons of peat waiting to be used for fuel or other purposes. *Peat has been an important fuel in Europe for hundreds of years.* If the decaying vegetable matter is exposed to the air and becomes alternately water-soaked and dry, it decomposes rather rapidly and forms **muck**. This is soft and spongy and does not show any trace of what it formerly was. It is usually mixed with sand and clay which have been washed over it.

When a small quantity of soil is viewed with the naked eye, it appears made up of small particles which look like little grains. They are tiny pieces of disintegrated and decayed rock which range in size from those that may be readily seen to those that are as fine as dust. Besides these rock particles, which are the *mineral matter*, there are also present pieces of roots, stems, leaves, etc., which are the *organic matter. In peat soil the organic matter predominates.* The rock particles that are large enough to be readily seen are called **sand**, while the very fine dustlike particles are termed **silt** and **clay**. The latter represents the very finest material. A magnifying glass is required to see the finest sand particles and to distinguish the silt from the clay. There are several important **classes of soils** that are commonly recognized. So when the sand particles are abundant, the soil is called a **sandy soil**. If the sand grains are not abundant, and the soil is quite floury when crushed, it is a **clay soil**. When there is a considerable proportion of sand present but also some clay, the soil is called a **loam**. The popular expressions *sandy, light sandy, loam, heavy clay loam,* and *clay,* all express the relative proportions of clay, silt, and sand in the soil. *Loams*

generally make good soils. They are usually fertile and easily worked. The vegetable matter in sands, loams, and clays is called humus, and **muck soils** are sometimes spoken of as **humus soils.** Real *humus is vegetable matter so completely decayed that one cannot tell what it was originally.*

Though thus made up of minute rock particles and tiny pieces of decayed roots and stems, the *soil nevertheless contains the substances necessary to make plants grow and develop to maturity.* When a seed is placed in soil of proper moisture content, it absorbs water and germinates. Firming the soil about the seed brings it into closer contact with the moisture and hastens the absorption of water. But *besides moisture, germinating seeds require oxygen,* and so while the soil must be in close contact with the seed and contain sufficient moisture, it must not be water-logged nor so closely packed around the seed as to exclude air, otherwise the seed will fail to germinate and will actually rot.

After germination has taken place a seed will continue to grow for some time at the expense of the reserve food which it contains; but sooner or later growth will cease unless certain chemical elements are available for absorption by the plant. *These so-called* **essential elements** *without which the plant cannot grow are nitrogen, phosphorus, sulphur, potassium, calcium, magnesium, iron, and hydrogen and oxygen in the form of water.* Plants indeed also contain other elements like sodium, chlorine, and silicon, but these are generally considered as non-essential, for plants have been grown without them. It should be clearly understood that *the chemical elements mentioned are absorbed by the rootlets of the plants in the form of salts that are in solution in the soil water.* Humus soils contain a great deal of carbon, but *the plant gets its carbon content from the carbon dioxide of the air through its leaves* (see Chapter IX), *and not from the soil through its*

roots. Some of the elements are used to build up the organic structure of the plant, while others remain in the plant as salts. So while hydrogen and oxygen unite to form water, hydrogen, oxygen, and carbon combine to form *starch, cellulose, sugar, fats, and oils* (see Chapter IX). Nitrogen, sulphur, and phosphorus together with carbon, hydrogen, and oxygen form the very complex substances called *proteins* (see Chapter IX), which are absolutely necessary for the growth of the plant.

Each crop taken from the soil removes a certain amount of the essential elements which the soil contains. The amounts of such elements thus removed have been carefully ascertained by chemical analyses. These results, together with the chemical analyses of the soils themselves, have shown that *most soils contain enough of the essential elements to furnish crops for hundreds of years.* Nevertheless, as will be further explained below, *the available supply of these elements is so limited that additional amounts must be added either as commercial fertilizers or as farm manures,* if crop production is to continue at a profit. Chemical analyses have also shown that an acre surface foot of soil may contain enough potassium to last 1500 years, and supplies of phosphorus sufficient for 200 to 500 years, of nitrogen for 100 to 300 years, and sulphur for 200 to 300 years. The amounts of plant food removed from an acre by some of our common crops are given in Table 8.

Besides the essential elements above mentioned, *plants require large quantities of water throughout their period of growth.* This water is necessary to dissolve the soluble salts in the soil and to convey them through the walls of the rootlets into the plant. Water is also needed to keep the cell walls of the leaf moist so that it may absorb carbon dioxide, and regulate the temperature of the plant; for *plants*

JEAN BAPTISTE BOUSSINGAULT, 1802–1887.

The first chemist who analyzed crops and manure and studied closely the relations of chemistry to farming.

TABLE 8

AMOUNT OF PLANT FOOD REMOVED FROM THE LAND BY VARIOUS CROPS, EXPRESSED IN POUNDS PER ACRE

KIND OF CROP	DRY WEIGHT OF CROP	N	P	S	K	Ca	Mg
	lb.	lb.	lb.	lb.	lb.	lb.	lb.
Wheat (grain), 30 bu.	1530	34	6.2	2.5	7.7	0.7	2.1
Wheat (straw) . .	2653	16	3.0	3.7	16.2	5.8	2.1
Total crop . .	4183	50	9.2	6.2	23.9	6.5	4.2
Barley (grain), 40 bu.	1747	35	7.0	2.6	8.0	0.8	2.4
Barley (straw) . .	2080	14	2.0	3.0	21.5	5.6	1.7
Total crop . .	3827	49	9.0	5.6	29.5	6.4	4.1
Corn (grain), 30 bu.	1500	28	4.3	2.5	5.4	0.3	2.0
Corn (stalks) . .	1877	15	3.5	2.2	24.7	7.5	3.3
Total crop . .	3377	43	7.8	4.7	30.1	7.8	5.3
Red clover hay . .	3762	98	10.8	6.1	69.2	63.9	16.9
Alfalfa hay . . .	9000	197	17.4	25.9	176.7	151.1	31.8
Turnips (roots) .	3126	61	9.7	23.1	89.6	18.1	3.4
Sugar beet (roots) .	4320	53	8.8	3.8	72.9	7.1	7.8
Potatoes (tubers) .	3360	46	9.4	1.1	64.5	2.4	3.7
Tobacco (leaf) . .	1800	65	3.5	6.4	73.8	57.5	15.0
Tobacco (stalk) .	3200	32	3.5	2.0	40.6	10.5	3.0
Total crop . .	5000	97	7.0	8.4	114.4	68.0	18.0

regulate their temperature by transpiration (*i.e.* exhalation of water vapor, from their leaves) just as animals keep their temperatures normal by means of perspiration from their skins. The amount of water thus required by a plant is very great. So, for example, *it takes from 250 to 300 pounds of water to produce a pound of dry matter in the corn plant, and in the clover plant from 500 to 600 pounds of water are necessary to finally obtain a pound of dry material.*

The **fertility of the soil** depends upon a number of factors.

In order to be fertile a soil must first of all contain the chemical elements that are essential for plant growth. But these elements must also be available to the rootlets of the plants in the form of suitable soluble compounds. Now since *nothing can pass into the roots except as it is dissolved in the soil water,* it is clear that plant food must always be present in solution in the water of the soil in sufficient quantity in order that crops may thrive. Furthermore, undesirable bacteria and other injurious organisms must not be present, that is to say, *the soil must be in proper sanitary condition.* Again, a sufficient quantity of moisture, containing carbon dioxide to enhance its solvent power, must be present in the soil, and the latter must be sufficiently loose to enable the roots to penetrate it and air to enter and circulate in it. At the same time soil should not be so porous as to dry out too rapidly. Sandy soils commonly present the latter defect. As has already been stated, *the solvent action of water upon the mineral matter of soils is rather slow,* even when the water is charged with carbon dioxide; and so, though a soil may contain a great abundance of the elements necessary for plant growth, this supply is often not available in sufficient quantity as needed. The decaying organic matter in the soil, to be sure, liberates carbon dioxide, which helps to increase the solvent power of the soil water, at the same time this rotted plant material furnishes to the soil in available soluble form the very elements that are essential for the growth of the crop. For these reasons *the presence of considerable amounts of decomposing vegetable matter is essential to the fertility of most soils.* But often the entire supply of plant food in a soil is insufficient, and so *it is necessary to fertilize the land,* that is, to add material to it which contains abundant plant food in readily soluble, *i.e.* available, form. This is accomplished by treating the soil with manure or commercial artificial fertilizers that contain the ingredients

which are lacking. Thus from the decomposition of mineral particles, from the remains of roots and other vegetable matter left on the soil, as well as from manures and fertilizers used, there gradually accumulates a considerable supply of material that serves as food for the season's crop, which as it is taken from the land in turn leaves its roots and stems to decay, and so, aided by additional manure and fertilizer, furnishes the material necessary for the next crop, and so on. But *there are always losses of plant food by erosion, the leaching of manure, and the sale of grain, sheep, cattle, hogs, milk, and other products from the farm.* These constantly cause a reduction of material for any succeeding year's crop. On the other hand, there may be very substantial gains made, especially of desirable compounds of nitrogen and carbon, by raising crops that absorb nitrogen from the atmosphere, a process which is accomplished by the action of bacteria growing on the roots. Clover, alfalfa, beans, peas, as well as other members of the legume family, are crops of this character. It is, of course, not always easy to estimate the extent of the losses which a soil has sustained, especially by erosion and waste; nor is it simple to ascertain the gains that have been made by growing legume crops.

Usually, *the organic matter and essential elements present in available form, together with the soil reaction, are* **the most prominent factors that are to be considered in crop production.** The humus produced by the decomposition of organic matter in the soil forms a coating around the soil grains. It helps to bind the soil grains together so that they are less readily blown about by winds. It also increases the moisture-retaining power of the soil, so that it can hold from two to three times its own weight of water. Again, in clay soils it lessens the tenacity with which the particles of the soil are held together, and this makes it easier to work these soils

properly. Humus is, moreover, a storehouse of nitrogen compounds, and consequently *humus soils are always rich in nitrogen.* By the slow action of bacteria, complex nitrogenous compounds in humus are successively broken down into the simple compounds, ammonia, nitrites, and nitrates. *The* **nitrates** *are the final and highest oxidation product.* They serve as a source of nitrogen for the growing plants, whose roots absorb the nitrates dissolved in the soil waters. The process by which ammonia is transformed to nitrates in soils is called **nitrification.** The amount of nitrates in the soil waters at any one time is very small, yet it is very important, for most plants take their supply of nitrogen from the soil in this form. *Plants utilize the nitrogen thus obtained in building up those very important and complex constituents, the proteins.*

The preparation of the available supply of nitrogen for plants by the bacteria of the soil forms one of the most interesting chapters in the story of the feeding of crops. When plants die, their roots, stalks, and leaves are returned to the soil, where their decomposition is brought about by molds, bacteria, and other low forms of life. Much plant material is fed to animals, but even then a large residue of unoxidized matter is returned to the soil as manure. This vegetable matter or animal refuse is attacked by putrefactive bacteria, which reduce the proteins to simpler compounds. Still other bacteria convert these into ammonia, which really represents the first step in the process of nitrification, *i.e.* nitrate making. Ammonia is changed to nitrites by the nitrite bacteria, and finally the nitrate bacteria effect the oxidation of nitrites to nitrates. The bacteria do not really produce the nitrites and nitrates, but they oxidize ammonia to nitrous and nitric acid. As these acids accumulate, they stop further action on part of the bacteria; hence the neces-

sity of having limestone or air-slaked lime present in the soil to neutralize the acids as they form. In this way *calcium nitrate is left in the soil as the final product of nitrification.* In the form of nitrates, the nitrogen then passes from the soil into the plant, where it is utilized in building up proteins. These are then returned to the soil direct, or they are first fed to animals, in which case the nitrogen is excreted in simpler compounds (urea, uric, and hippuric acids) in the urine and more complex ones in the feces. When returned to the soil, *all of these compounds are first converted to ammonia and finally again to nitrates, and thus the* **nitrogen cycle** *completes itself over and over again as long as the sun furnishes the necessary energy for plant growth.*

FIG. 76. — Nodules on the roots of a soy bean.

While plants in general do not get their nitrogen supply direct from the air, it has already been mentioned that legumes are able to store up nitrogen from the air in little nodules on their roots, with the aid of certain bacteria. So, for example, the nodules on the roots of red clover, beans, and cowpeas are a matter of common observation. In these nodules are bacteria which take the nitrogen direct from the air, and then give it over to the plant in the form of nitrogenous compounds. This process of converting the nitrogen of the air into nitrogenous compounds is commonly

spoken of as the **fixation of nitrogen.** The plant in turn gives the bacteria such food as sugar to live on. When the plant dies, there is left in these nodules on the roots a considerable supply of nitrogen, which upon subsequent nitrification becomes available for the next crop Thus it is that *the growing of a good crop of cowpeas or soy beans will usually leave enough available nitrogen in the soil for a good crop of cotton or corn. Consequently raising a crop of a legume in a rotation is one of the most important methods of maintaining the available nitrogen supply of the soil.*

Soils are spoken of as acid, alkaline, or neutral according to their *reaction toward litmus paper.* **Acid soils** are unsuitable for the successful growth of legumes, and it is therefore usually desirable to correct such acidity by an application of fresh lime, ground limestone, or marl. It is better to use old, thoroughly air-slaked lime, rather than fresh quicklime, for the latter has caustic properties. A ton of finely pulverized limestone per acre once in five years is commonly not too much.

The **amount of phosphorus** in most soils is generally small, being from 0.02 to 0.1 per cent. *All crops take phosphorus from the soil, in which this element is present as phosphates.* Plants store up phosphorus, especially in their seeds. The supply of phosphates in the soil can consequently be maintained by lessening the sale of seed, or crops, or by adding suitable phosphorus containing fertilizers. If a farmer buys plentifully of feeds, especially bran, for his stock and sells only milk and animals, he may not need to purchase phosphorus in the form of commercial fertilizer. Indeed, *the soils of farms of dairy states, where milk and cream are sold and feeds are purchased, may actually show a gain of phosphorus.* Raw **rock phosphate**, also called " **floats**," is a cheap commercial fertilizer for soils deficient in phosphorus. When used

with organic material, although rather slow in becoming available on account of the fact that its phosphates do not dissolve very readily, it nevertheless gives good results. **Superphosphate**, or **acid phosphate** (which, as has been stated in Chapter VII, is made by treating bones or phosphate rock with sulphuric acid), dissolves far more readily and hence is immediately available to plants. It should be used when a supply of organic matter is not at hand.

There is usually a sufficient amount of **available potassium in soils.** *This is especially true of clay soils, which are rich in feldspars.* And so when these soils are furnished with plenty of organic matter that gives off carbon dioxide as it decomposes, thereby increasing the solvent action of the soil water on the feldspathic constituents, an ample supply of potassium is furnished to the rootlets of the growing crop. *Some soils, like sands and marshes, are deficient in potassium.* To these it should be added at the rate of fifty to one hundred pounds of potassium chloride or potassium sulphate per acre.

Although **the quantity of sulphur in soils** is usually low, being about as low as that of phosphorus, it is replenished to a considerable extent during the year by rains. These carry sulphur dioxide down in solution from the atmosphere. The sulphur dioxide has come into the air mainly from the burning of coal. *For some crops which need much sulphur, like onions, cabbage, turnips, and rutabagas, additional sulphur in the form of gypsum, calcium sulphate, is advantageous.* From 100 to 200 pounds per acre is a good application.

It has already been stated that **the soil must be in proper mechanical condition** so that it will readily hold the proper amount of water, permit air to circulate in it, and allow the roots to penetrate it. *Proper* **tilth of the soil**, *that is to say, proper mechanical working of the soil with the plow, spade,*

hoe, harrow, roller, etc., is of great importance. While it is desirable that the particles of the soil be sufficiently fine, it must also be remembered that they may be so fine and closely compacted together that neither air nor root hairs can gain entrance. This condition is just as unfavorable as a coarse, lumpy soil. In either case the water storage capacity is decreased. From soils that are too loose and porous, moisture drains away, and the air circulates so freely in them that they dry out too rapidly. It has already been stated that this is a common defect of sandy soils. By the force of capillarity, or **capillary attraction**, water rises in tubes of small bore or between particles of solid substances. Fine-grained soils, of which clay soils are an excellent example, have small pores or spaces between their particles, and hence water will rise from below and be nearer to the surface in these soils than in those of coarser grain. *It is indeed possible to aid the upward capillary movement of water* so that in a dry season, when the seeds have been planted, moisture will be drawn to them from below. This is accomplished by rolling the ground, for rolling forces the particles of soil closer together and thus a stronger upward movement of water by capillarity results. *Much water that has thus been drawn up by capillarity is lost by evaporation at the surface of the soil.* This loss can be greatly reduced by covering the surface with a protecting layer or soil **mulch.** So if the surface of the soil to the depth of two or three inches is thoroughly stirred on a dry, windy day, the dry layer that is thus formed becomes an effective barrier against large losses of moisture from below. In regions where rainfall is light this practice is of the very greatest importance. It is upon this principle that so-called **dry farming** is based.

In speaking of **soil waters**, the terms *gravitational water, capillary water,* and *hydroscopic water* are often used. Their

meaning will be clear from the following experiment. If the neck of a funnel be stoppered, and the funnel be filled with soil, and then water poured on till the funnel is brim full, *the soil will be in a saturated condition.* If now the stopper is removed, a considerable portion of the water will drain off; this is called **gravitational water,** for it runs off because of the action of the force of gravitation. If the lower end of the funnel is now stoppered again and the whole allowed to stand, with the upper surface exposed, a large part of the moisture retained by the soil will rise up to replace that which is lost at the surface by gradual evaporation. This upward movement of the water in the soil is due to capillarity, and the water held and moved in this way is termed **capillary water.** If the funnel is thus allowed to remain exposed to the air for several days or weeks, the soil it contains will become dry. But if a portion of this apparently dry soil be weighed out and then heated at the temperature of a boiling water bath, 100° C., it loses still more moisture. The water thus expelled is called **hydroscopic water.** It is obvious that it is held more tightly by the soil particles than the capillary water.

The removal of gravitational water by proper natural or artificial underdraining is necessary because, if it remains in the soil, the latter becomes water-logged, and so displaces the air which is necessary in the soil for the growing roots. **Drainage** *lowers the water level in the soil, causing roots to penetrate deeper for moisture, and consequently their feeding ground is increased.* This helps them to withstand drought. The amount of moisture in a soil must be taken into consideration in determining whether it is an opportune time to work the soil or not. So, for instance, if heavy clay soils are worked when wet, the particles are forced close together and "**puddling**" is the result. This is a brickmaker's term. In making

brick the first effort is to destroy the granular texture of the clay, which is accomplished by wetting and working it. If this is done with clay soil, it obviously will be left very compact and hard when it dries. Such a condition, once brought

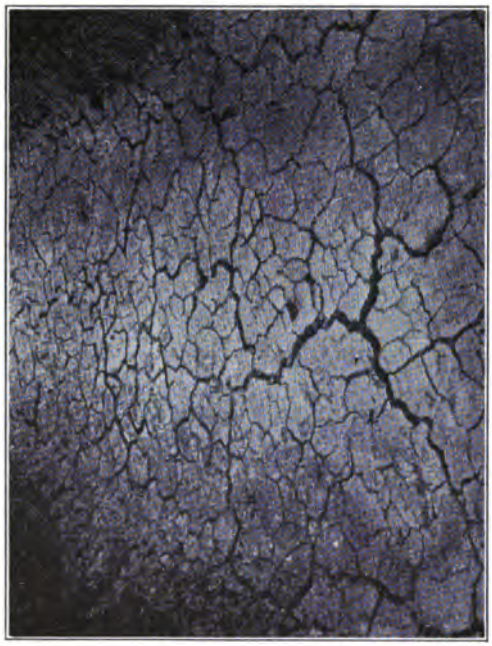


FIG. 77. — Cracked soil. Large losses of water will result from such a condition.

about by working the clay soil when too wet, may last for a number of years, and so *it is wise for the farmer not to get his clay fields puddled.* When land breaks up cloddy, it is desirable that it be plowed long before planting in order that freezing, thawing, and weathering may break up the clods and make them loose and mellow. This requires time, but for best results long exposure to the weather will pay. The

effect of humus on fine-grained soils, such as clays, is to divide the particles and so lessen the tendency to puddle. In sandy soils, humus tends to cement the grains of sand together more firmly, a result that could not be secured by mere films of water around such coarse particles.

The **temperature of soils** is affected by their power to absorb heat from the rays of the sun. The readiness with which heat is thus absorbed is greatly influenced by the **color of the soil**. Thus dark-colored soils absorb heat readily and so warm up relatively rapidly. They are consequently called *warm soils*. On the other hand, light-colored soils do not absorb heat so readily as dark-colored ones of equal moisture content, and they consequently do not warm so quickly. They are called *cold soils*. It must be borne in mind, however, that *the amount of moisture in the soil is the most important factor in determining the rate at which it will warm up*, for it takes large amounts of heat to evaporate water, and so *dry soils always warm up more rapidly than wet ones, regardless of their color or composition*. Barefooted boys well know that dry sand and fine road dust become warm more quickly than a wet soil. Indeed, the evaporation of water from the soil cools the latter, just as the evaporation of sweat cools the animal body. *Dark-colored, sandy soils warm up earliest in spring* because the gravitational water soon drains out of them and they readily absorb heat from the sun's rays. *Such soils, then, are most suitable for market gardening or early spring crops.*

QUESTIONS

1. From what is soil made, and what are the agencies active in its formation? What is a glacier?

2. How are peat bogs formed?

3. What is the difference between sedentary and transported soils?

4. What is the difference between a sandy soil and a clay soil? What is humus?

5. Name the essential elements needed for crop production.

6. About how much phosphorus is there in a soil, and how many crops of wheat would it produce?

7. How much water is required to make a pound of dry matter in a corn plant?

8. Why is it necessary to have plenty of organic matter in the soil?

9. How would you maintain the nitrogen content of soils?

10. What may the reaction of the soil be, and what should be the proper reaction for the growth of legumes?

11. What is meant by the puddling of clay? What is the influence of humus on the water-holding power of a soil? Distinguish between capillary and gravitational water.

12. How does drainage help a plant in a season of drought?

CHAPTER XV

COMMERCIAL FERTILIZERS

COMMERCIAL fertilizers are manufactured plant foods. They are usually kept for sale at warehouses and seed stores. About $100,000,000 is spent annually in the purchase of fertilizers in the United States and probably one-half of this money is thrown away. This is not an argument against the use of commercial fertilizers, but it does mean that they should be purchased with proper judgment, and not used at all until trials have shown that they are actually necessary.

Experimental investigations and common experience have shown that material increases in the yield of the land have in most cases been obtained by adding to the soil but three of the essential elements necessary for plant growth; namely, *nitrogen, phosphorus,* and *potassium.* These elements, it is to be recalled, *are not applied in the free or uncombined state, but in the form of compounds, usually salts, that will dissolve in the soil waters, so that they can be absorbed by the rootlets of the growing crop.* In consequence of the results of experience, then, commercial fertilizers as placed on the market to-day contain as their essential ingredients only nitrogen, phosphorus, and potassium in the form of mixtures of salts or other compounds of these elements; and thus it is that *the nitrogen, phosphorus, and potassium in a commercial fertilizer are the only materials which give that fertilizer value.* It should be remembered, however, that equally good results are sometimes secured by applying to the land dressings that

(*A*)

(*B*)

FIG. 78. — Effect of fertilizers on a crop of potatoes. (*A*) Complete fertilizer used. (*B*) No fertilizer used.

do not contain any of the three elements just mentioned. So, for example, lime or limestone and gypsum are frequently beneficial. *The benefit derived from lime is usually due to its basic properties, which cause a neutral or alkaline reaction of the soil.* Gypsum, on the other hand, acts as a source of sulphur, and also produces physical changes in the soil.

Commercial fertilizers are made from a few basal materials which are articles of commerce. A so-called *complete fertilizer* consists of two or more such basal materials mixed together so as to give the desired percentage content of nitrogen, phosphorus, and potassium. *These basal materials are derived* (1) *from inorganic or mineral matter, and* (2) *from organic matter, that is to say, from animal or vegetable material.* So, for example, as **sources of nitrogen**, the inorganic salts, ammonium sulphate and sodium nitrate, and organic materials like dried blood, meat scraps, tankage, fish scrap, and cotton-seed meal are in common use. As **sources of phosphorus**, phosphate rock and basic slag, which are inorganic, and bones and guano, which are organic, are utilized. Finally, as **sources of potassium**, potassium chloride, potassium sulphate, and kainite, all of which come from the Stassfurt salt deposits, are employed. All of these materials will now be considered more in detail.

Among the **nitrogenous fertilizers** ammonium sulphate, $(NH_4)_2SO_4$, also popularly called " sulphate of ammonia," is used to a considerable extent. This salt has already been mentioned in Chapter IV. It is a by-product of the manufacture of illuminating gas from coal, also of the dry distillation of bones in the manufacture of bone black. **Ammonium sulphate** *is a very concentrated fertilizer, containing 20 per cent of nitrogen.* It is copiously soluble in water, does not readily leach out of the soil, and very quickly undergoes

nitrification, that is, conversion into nitrates. Sodium nitrate, $NaNO_3$, also called **nitrate of soda**, has been described in Chapter X. It occurs as *Chili saltpeter* in extensive deposits in the rainless districts of western South America. The natural material from the saltpeter beds contains a large amount of common salt. This, however, is practically removed from the product before it is placed on the market, so that as sold it is a *crude nitrate of soda containing from 15 to 16 per cent of nitrogen.* Nitrate of soda is readily soluble in water and easily leached from the soil. *It should consequently be applied just before planting or after the plant has possession of the soil.* Fifty to one hundred pounds per acre is a proper application. Dried blood, meat scraps, and tankage come from the slaughterhouses. **Dried blood** is simply blood of the slaughtered animals dried and ground. Like all organic nitrogenous fertilizers, it is insoluble in water; but it ferments readily in the soil and yields its nitrogen in soluble form to the growing plant. *Dried blood contains from 13 to 14 per cent of nitrogen,* while **meat scraps** contain variable amounts of nitrogen, usually from 4 to 10 per cent. The value of fish as a fertilizer was known to our American Indians. Squanto, an Indian, taught the New England settlers that they could increase the yield of corn by putting a fish under each hill. This gave the grain the two elements it needed most, phosphorus and nitrogen. To-day **fish fertilizers** are made in large amounts from rough fish and fish offal on the Atlantic coast and the Great Lakes. **Cotton-seed meal** is obtained by removing the hulls and oil from cotton seed. The residue is then ground and put on the market. It is extensively used as a fertilizer in the south. There are still other nitrogenous fertilizers like **waste leather, hoof and horn meal, hair** from the slaughterhouses, and **wool waste** from the woolen mills. The chemical nature

of these has already been described in Chapters IX and XIII. All of these materials decompose so slowly in the soil that many states have passed laws prohibiting their sale as fertilizers. *In ordinary farming it is seldom profitable to purchase nitrogenous fertilizers,* for the nitrogen supply of the soil can be maintained by means of farm manure and the proper use of legume crops in rotation, as already stated in Chapter XIV. Such crops as the clovers, peas, and beans are aids to the farmer as nitrogen gatherers, especially if he occasionally plows under a goodly part of the plant. *In intensive farming, like market gardening, it will be necessary to make liberal use of nitrogenous fertilizers of commercial origin.*

Since three-fourths of the phosphorus absorbed from the soil is deposited in the grain of the crop and therefore ordinarily sold from the farm, it is clear that **phosphatic fertilizers** *are of fundamental importance.* In the soil, phosphorus occurs as the phosphates of calcium, magnesium, and iron. As already pointed out in Chapter VII, phosphates are salts of phosphoric acid, H_3PO_4. *In all the important commercial phosphate fertilizers, including phosphate rock, bones, basic slag, and guano, phosphorus is present as phosphates formed by the combination of phosphoric acid,* H_3PO_4, *with lime,* CaO. These are phosphates of calcium, commonly spoken of as phosphates of lime. With lime, phosphoric acid forms three important compounds: (1) *insoluble phosphate of lime,* which is tricalcium phosphate, $Ca_3(PO_4)_2$; (2) *soluble phosphate of lime,* which is monocalcium phosphate, $CaH_4(PO_4)_2$; and (3) *reverted phosphate of lime,* which is dicalcium phosphate, $Ca_2H_2(PO_4)_2$. The first of these, as its name indicates, is almost insoluble in water and hence is not readily available to plants. Ground bone, guano, and floats contain their phosphorus in this form. **Ground bone,** or bone meal, is a product of the packing houses,

glue factories, and soap works, the raw material being the bones of farm animals. The bones are either ground directly or after they have been steamed and dried. In the latter case they are called **steamed bone**. *Raw bone contains 2.5 per cent of nitrogen and 11 per cent of phosphorus.* It is more effective when finely ground than when coarse; furthermore, raw bone decomposes more slowly in the soil than steamed bone, because the fat has been removed from the latter. **Guano**, a highly prized but comparatively rare fertilizer, consists of the excrements and remains of sea fowls, which have accumulated in certain localities along the west coast of South America. *It contains nitrogen, potassium, and phosphorus, the latter amounting to from 4 to 8 per cent.* The phosphorus in guano has come from the skeletons of of the birds. **Floats** is the term applied to the **finely ground**

FIG. 79. — Phosphate-rock mining in Florida.

crude phosphate rock found in Tennessee, the Carolinas, Florida, Georgia, and other Southern states. It has recently also been discovered in considerable beds in Utah and

Wyoming. *It contains from 11 to 13 per cent of phosphorus and is the chief source and also the cheapest source of phosphorus supply on the market.* Recent investigations have shown that when used with sufficient organic matter, such as farm manure or green manure, ground phosphate rock has high fertilizing power, for the organic matter on decomposing furnishes carbonic acid, which aids in dissolving the tricalcium phosphate, thus making it available. **Soluble phosphate of lime** is also known as *"one lime phosphate," acid phosphate, acidulated rock, superphosphate, etc.* It is made by treating bones or phosphate rock with sulphuric acid, as stated in Chapter VII, where the chemical equation explaining its formation is given. Because of its solubility in water, *this fertilizer contains the phosphorus in the most available form for direct use by the plant. A good sample contains about 7 per cent of phosphorus,* or only about half as much as a good sample of rock phosphate. Furthermore, a ton of the material contains about half a ton of gypsum. The latter is formed simultaneously when the sulphuric acid acts on tricalcium phosphate. Though readily soluble in water superphosphate is not leached out of the soil, for the lime and iron in the soil can form insoluble phosphates with it. In the process of making acid phosphate the whole of the insoluble phosphate is generally not acted upon. The tricalcium phosphate thus remaining when left with large quantities of the monocalcium phosphate forms dicalcium phosphate, thus:

$Ca_3(PO_4)_2$	+	$CaH_4(PO_4)_2$	=	$2\ Ca_2H_2(PO_4)_2$
tricalcium phosphate from bone or phosphate rock		monocalcium phosphate, "acid phosphate"		dicalcium phosphate "reverted phosphate"

It is clear, then, that the dicalcium phosphate formed is less acid than the monocalcium phosphate, *i.e.* the latter has

"reverted" somewhat. Hence it is known as "**reverted phosphate.**" Though almost insoluble in water, it dissolves in water containing carbon dioxide and hence is readily assimilated by plants. As stated in Chapter XI, a phosphate fertilizer is produced in the manufacture of steel. This material is popularly called **basic slag,** its composition being about $Ca_5P_2SiO_{12}$. Though it was formerly thrown away, its value as a fertilizer is now duly recognized. *It contains from 6.5 to 8.5 per cent of phosphorus* in a somewhat different form from that in the other phosphates just described, as the formula given indicates. *It should be finely ground,* for it is difficultly soluble in the soil water. *The fact that it is rich in lime makes it specially desirable for use on acid soils.* Most of the basic slag on the market is imported from England, but it is not unlikely that more will soon be made in the United States, as iron ores of high phosphate content are used in the iron and steel industry.

The **potash fertilizers** are commonly considered as of less importance than either the nitrogenous or phosphatic fertilizers, for potassium compounds are usually more abundant in the soil than either nitrogen-bearing compounds or phosphates. Moreover, *though most crops remove more potassium than phosphorus, the former element is more liable to be returned to the soil than the latter.* Potassium is more abundant in the stems and leaves of plants, and as these are the parts that are generally returned to the land in the form of manure, the supply of potassium is constantly being replenished. On the other hand, *potassium is a very necessary constituent of fertilizers intended for light sandy soils, for marsh land, and for fields that are to produce crops which consume large quantities of potassium, like potatoes, tobacco, and fleshy roots.* Potash fertilizers can be applied in the fall on heavy clay soils, or in the spring on sandy soils. As stated in Chapter XIV, clays

generally do not need potash fertilizers if well supplied with organic matter. The potassium salts used as fertilizers have already been described in Chapter X. They come from the Stassfurt mines. So the chloride KCl, also called the **muriate of potash**, *contains about 47 per cent of potassium*, which is combined with chlorine, as the formula indicates. This salt can be used on all crops except a few like tobaccco, sugar beets, and potatoes, which appear to be injured by the chlorine present. Potassium sulphate, K_2SO_4, also called **sulphate of potash**, *contains about 44 per cent of potassium*, and is of special value for those crops that are injured by the chloride, as just mentioned. **Kainite** (see Chapter X), *the most common of the Stassfurt fertilizers*, is essentially a double salt of potassium chloride and magnesium sulphate. *It contains from 10 to 12 per cent of potassium* as chloride and sulphate. It is thus a low-grade potash fertilizer; and while cheap per ton, it really costs more than the muriate or sulphate, for these are so much richer in potassium. For many years **wood ashes** were the sole source of potassium for fertilizing purposes, but since the introduction of the Stassfurt salts, ashes appear much less on the market. When unleached, *ashes contain from 2 to 7 per cent of potassium and hence are valuable.* Furthermore, ashes contain the potassium as carbonate, K_2CO_3, and also carbonate of lime, $CaCO_3$, both of which are excellent for neutralizing, that is, "correcting," soil acidity.

A number of substances like lime, gypsum, and common salt are beneficial to the land under certain conditions, but they are generally not considered as important sources of plant food, and so are frequently termed **indirect fertilizers**. *There are but few if any soils that do not contain enough lime to supply the needs of the plant, and so* **the value of lime** *as a fertilizer is indirect.* It corrects acidity, acts upon other compounds

so as to liberate soluble potassium salts, and helps bacterial life in the soil. By keeping the soil neutral, it creates a proper home for many microörganisms which help to make plant food available. Lime for agricultural purposes is put on the market as caustic or burned lime, CaO; air-slaked lime or carbonate of lime, $CaCO_3$; ground limestone rock; ground clam or oyster shells, $CaCO_3$; and also as beet sugar refuse. The chemical nature of lime, limestone, or calcium carbonate, and gypsum has already been sufficiently described in Chapters IX and X. **Air-slaked lime** is quicklime which has again been converted to the carbonate, $CaCO_3$, by exposure to the air. When caustic lime is used on the land it must be applied with care, as already stated in Chapter XIV. **Finely ground limestone** is coming into general favor, and when it can be obtained at a sufficiently low cost, it is undoubtedly the safest form in which to apply lime, especially by the inexperienced. Lime should be spread over the surface and then incorporated with the upper two to four inches of the soil. *The clovers and legumes in general require more lime than do the cereals, for they are much more sensitive to the acidity of the soil.* A ton per acre every four or five years is a good application of limestone. Gypsum, **land plaster,** $CaSO_4 . 2\ H_2O$, has given excellent results with clovers and legume plants. It is also a cheap source of sulphur, which is low in amount in most soils, and is freely used by plants of high protein content as well as by such crops as cabbage, kohlrabi, turnips, rutabagas, onions, garlic, and horse-radish. **Common salt,** NaCl, has sometimes been used as a manure, but it is not supposed to furnish necessary plant food, and while it has given beneficial results at times, often it has failed to do so.

Mixed fertilizers are very common on the market. The fertilizer manufacturer very generally does the mixing of the

basal materials already described, and sells them to the farmer as complete fertilizers. Mixed fertilizers are indiscriminately recommended for general use and all sorts of startling claims are made for them by the manufacturers. They are offered as universal fertilizers, irrespective of the plant food contained in the given soil or the needs of the crop to be grown. The " corn special " of one manufacturer may be like the " tobacco special " of another firm. *The farmer must study the real needs of his fields, and when he has done this, he will prefer to buy the basal fertilizing materials of definite and known composition and apply the proportion that is best adapted to his needs, rather than buy mixed goods.* It is wise to practice **home mixing**. The difference in the cost of complete fertilizers and the basal materials per pound of plant food is considerable. It has been estimated to be from eight to ten dollars per ton. On the farms of the eastern United States it is becoming common practice to **buy the basal materials and mix them at home**.

The proper selection of commercial fertilizers is obviously of great importance. *It is impossible to give definite directions as to kinds and quantities of fertilizers for different crops, because soils differ greatly in their content of plant food.* But by carefully noting the growing crop, some valuable suggestions may be gained as to which fertilizers are most needed. So a crop with a deep green color, well-developed leaf, and luxuriant growth is *not suffering from a lack of nitrogen or potassium.* A rank excessive growth of leaf and stem, but imperfect bud and flower development, and small undersized grain, denotes *plenty of nitrogen and potassium,* but a *lack of phosphorus.* These suggestions are helps, but it is not always safe to depend upon them entirely. *Field experiments are necessary.* To determine with any degree of certainty what particular fertilizer would be helpful, *the*

farmer must conduct some experiments himself. This can be done by carefully marking off certain portions of the field of uniform soil into plots of definite size, and then using different fertilizing materials on these plots. A plot one rod wide and eight rods long contains one-twentieth of an acre, and is of convenient size for experiment. Careful notes should be kept during the growing season, and the weights of the harvested crops should be recorded. This gives definite information, especially if continued over several years. When it has thus been established what basal materials are most effective, then the cheapest sources of these materials should be selected. The diagram given below indicates the arrangement of the plots and the amounts of materials to apply. In addition to the fertilizer, a part of the plots may be limed. This will give evidence as to the need of correcting the acidity of the soil.

DIAGRAM

Plot 1. None.	Plot 8. N and S (Blood and Sulphate).
Plot 2. N (Dried Blood, 30 lb.).	Plot 9. P and S (Bone and Sulphate).
Plot 3. P (Steamed Bone Meal, 10 lb.).	Plot 10. P and K (Bone and Chloride).
Plot 4. S (Gypsum, 10 lb.).	Plot 11. N, P, and S (Blood, Bone, and Sulphate).
Plot 5. K (Muriate of Potash, 10 lb.).	Plot 12. N, P, S, and K (Blood, Bone, Sulphate, and Chloride).
Plot 6. None.	Plot 13. None.
Plot 7. N and P (Blood and Bone).	

No definite rule can be given as to the amount of fertilizer to apply, for soils vary in their original content of plant

food; and then the kind of crop to be grown is also a factor. *Five hundred pounds per acre on soils already fertile is a heavy application for ordinary farm crops* and will only give profitable returns on fertile soils when used in intensive farming. *It is better to make lighter applications of from 100 to 200 pounds of any one basal material per acre,* and to vary the amount from year to year, when experience has shown that economical returns can be expected from heavier applications.

FIG. 80. — A bag for commercial fertilizer with the analysis printed on it.

To protect the farmer against fraud, many state **fertilizer laws** require that the **guaranteed analysis** or formula of a fertilizer should be placed on the bag containing it. Most of such legends are so complicated as to confuse and are of no more value than so many Greek letters. A sample of such a label is given on page 227.

The *things that are essential in the statement* on page 227 are the amounts of "nitrogen," "available phosphoric acid," and "potash." *The rest is superfluous and only confusing.* The label means that there is contained in this fertilizer 8 to 10 per cent of available P_2O_5; but this is erroneously called

GENERAL CROP BRAND

GUARANTEED ANALYSIS

	PER CENT
Nitrogen	0.82 — 1.65
Ammonia	1.0 — 2.0
Available phos. acid	8.0 — 10.0
Equal bone phos.	17.0 — 21.0
Total phos. acid	10.0 — 12.0
Potash sulphate	11.0 — 13.0
Potash	6.0 — 7.0

phosphoric acid (see Chapter VII). It would be better if the label stated the real fact, which is that the fertilizer contains 3.4 to 4.4 per cent of available phosphorus, P, for this is the equivalent of the statement "8 to 10 per cent available phosphoric acid." Furthermore, potassium is expressed on the label as potash, meaning K_2O. The 6 to 7 per cent of potash are, however, equivalent to 4.9 to 5.8 per cent of potassium. *In expressing the analysis of fertilizers, the percentages of the* **elements** *nitrogen, phosphorus, and potassium ought to be given, because these represent the things of actual worth.* This simple and intelligible system of expressing the results of a fertilizer analysis is known as the **elementary system**, and it ought to be adopted universally.

QUESTIONS

1. What is a commercial fertilizer, and what elements are the main ones supplied in a complete fertilizer?

2. Name two kinds of nitrogenous fertilizers. How much nitrogen would there be in 200 pounds of Chili saltpeter? In 100 pounds of ammonium sulphate? How do these amounts compare with the requirements of 100 bushels of corn?

3. When should nitrate of soda be applied to the soil?

4. What is the objection to the use of scrap leather as a source of nitrogen for fertilizers?

5. Name three kinds of phosphate fertilizers. How does an acid phosphate differ from floats?

6. What does basic slag come from? On what kind of soils is it doubly useful?

7. Where do potash salts mainly come from? What compounds of value are contained in wood ashes?

8. In what form is it best to add lime to a soil, and why should it be added? Of what use is gypsum when added to soils?

9. How are fertilizers commonly sold? What is meant by home mixing?

10. How can the fertilizer requirements of a soil be best determined?

11. A mixed fertilizer contains 2 per cent of nitrogen, 2 per cent of phosphorus, and 8 per cent of potassium. How many pounds of ammonium sulphate, bone meal, and kainite would be needed to make up a ton of such a mixture?

CHAPTER XVI

FARM MANURE

Of all fertilizers, farm manure is the oldest and still the most popular. It consists of the liquid and solid excreta of the farm stock, plus the bedding employed. Early Roman writers called attention to the fact that the application of the excreta of farm animals resulted in increased crop production, and from that time to the present day the majority of farmers have placed their reliance on such manures for maintaining the fertility of the land.

A well-kept manure heap may safely be taken as one of the surest indications of thrift and success in farming. The **fertilizer value of the manure** produced each year by the different classes of farm animals in the United States aggregates a total of $2,225,700,000. This estimation is based on the value usually given to phosphorus (P), potassium (K), and nitrogen (N) in commercial fertilizers, which are respectively 10 cents, 6 cents, and 15 cents per pound for the materials in the order named. *It is safe to say that the manure made on the farms of Wisconsin has a value of $75,000,000 per year, and that through improper handling one-third of the value is lost.*

The manure produced by the various classes of farm animals differs greatly in its composition and physical properties. The manure from the pig and cow (including both the solid and liquid excrement) is very high in water, containing at least 75 per cent. Because it is so moist, it ferments very slowly when piled, and so it has been called a **cold manure.**

On the other hand, the combined excreta from the horse and sheep have a lower water content, about 65 to 70 per cent. These ferment more rapidly and are called **hot manures.** While the composition of manure varies greatly, *a ton of mixed manure of average composition generally contains approximately 2 pounds of phosphorus* (P), *10 pounds of nitrogen* (N), *and 8 pounds of potassium* (K).

The composition and value of manure *is influenced by three things*: (1) *the nature of the ration fed the animal;* (2) *the kind and age of the animal;* (3) *the kind and amount of the bedding used.* The manure from animals that are fed on rich foods, like clover, alfalfa, bran, and cottonseed meal is better than that from animals fed on timothy hay, straw, corn stover, and other feeds not rich in nitrogen and phosphorus, and to which little or no grain has been added. Furthermore, the manure from young and growing animals, as calves and colts, is not so valuable as that from older animals, and especially animals that are fattening. *Young growing animals use up more of the nitrogen, phosphorus, and potassium in the food for their own bodies to make blood, bone, and flesh than do older animals. Mature animals that are poor in flesh, but are fattening, produce a rich manure, for the production of fat does not require the retention of nitrogen, phosphorus, and potassium, as does the building of bone and muscle.* A milch cow will not produce as rich manure as a fattening steer on the same ration. This is because some of the fertilizer constituents go over into the milk and so are not returned in the excreta. *There is in 5000 pounds of milk fertilizing material equal in value to about $4.90.* If the milk is sold, this is lost to the farm. If the milk is skimmed and the butter is sold, practically no loss occurs, for the nitrogen, potassium, and phosphorus are in the skimmed milk. *The fertilizer value of 500 pounds of butter is only about 10 cents.* This shows how

a dairy or animal farm has the advantage over a grain farm in helping to keep up the fertility on the land. The bedding employed will alter the composition of the final manure. *The richer the bedding, the richer the manure;* but the materials commonly used, such as the straws and sawdust, are low in fertilizer ingredients. If peat is used, it will increase the nitrogen content of the manure considerably.

The great value of farm manure is not appreciated to its fullest extent. A large proportion of our farmers seem to have little realization of the immense loss incurred through the waste of this important product of the farm. They begrudge the time and labor required to remove it from the farm and feeding lot, and it is not uncommon to see the purchase of commercial fertilizers and the waste of farm manure going on at the same time and on the same farm. Barns are often built on steep hillsides or the banks of running streams, which practice insures a large waste. *The* **value of manure may be increased** *in two ways:* (1) *by the selection and use of feeds that are rich in fertilizing materials, or* (2) *by purchasing suitable commercial plant food and adding it to the manure.* The first method means that the farmer must purchase some of the mill by-products now offered for sale, and feed them to his stock. A successful stock man usually finds it profitable to thus reënforce his home-grown grains with one or more of the various mill by-products. *A farmer who is buying large amounts of concentrates is increasing the fertility of his land,* **provided he is saving the manure.** **Bran** is a concentrate that is rich in phosphorus and potassium, while **oil meal** and **gluten feed** are rich in nitrogen.

The second method of enhancing the value of manure is practiced in systems of animal husbandry as well as in systems of mixed farming. Experiments have shown that on certain soils, reënforcement of the manure with either kainite,

gypsum, floats, or acid phosphate results in profitable increase in crops. *The greatest profit, however, comes from the use of either acid phosphate or floats.* Any of these materials can be applied at the rate of 40 pounds per ton of manure. To secure the greatest profit, *the manures for most soils should probably be reënforced with floats and gypsum,* for these materials furnish additional phosphorus and sulphur respectively, of which elements there is often a lack.

Farm manure is a perishable product and must be handled with intelligence to obtain the greatest value from it. Doubtless, as manure is handled on most farms, only about one-half of its worth is realized. The greatest loss is through the waste of the liquid excrement by not using sufficient bedding to absorb it. The boring of holes in the floor for the express purpose of allowing the urine to run off as rapidly as possible is by no means an uncommon practice. *Pound for pound the liquid excrement is more valuable than the solid.* It contains more than one-half of the nitrogen, most of the potassium, and but a trace of the phosphorus excreted by the animal. Most of the phosphorus is found in the solid droppings. Another fact of great importance in this connection is that the plant food in the urine is soluble in water and consequently more available to plants than that in the solid dung. The solid excrement consists in part of the indigestible portions of the food, and before its nitrogen can become available to plants, it must undergo decomposition and decay. Besides the losses due to incomplete absorption of the urine by the bedding, *manure suffers heavily from leaching by rains.* Often animals are confined in yards and no effort at all is made to save the droppings, and the result is that the rains wash them away. Very often when stables are cleaned out, the manure is thrown under the eaves of the barn or shed and the rain from the roof soon saturates the pile and a dark

liquid begins to run away. This contains much plant food, and it is usually lost by draining into a stream or soaking into the ground where no plants are to grow. Then again, it often happens that manure is thrown into large piles that are allowed to lie exposed to the rain and weather for months. Under such conditions it gets "warm" and ferments.

FIG. 81. — An only too common way of wasting farm manure.

This fermentation is caused by bacteria, and results in the production of ammonia (NH_3), which contains nitrogen. The ammonia gas escapes into the air and is lost. Many a boy on the farm has suffered with smarting eyes when removing the accumulated manure from the horse stable. *Leaching and fermenting cause great losses to manure, which can largely be prevented.* Thus if the manure pile be made so compact that the air cannot penetrate it, the most destructive bacteria cannot live in it and hence hot fermentation is prevented. Other bacteria will be active in the pile and cause

rotting; but this is not harmful, being on the contrary a process greatly desired in the production of manure for market gardening. *The secret of keeping the hot fermentation down is to keep the manure heap compact and moist. This allows the pile to rot, but the loss of nitrogen as ammonia will be small.*

It is better to allow the manure to accumulate in the stable and be trampled underfoot by the animals, providing plenty of bedding is used, than to throw it out in a pile to ferment and leach. The tramping in the stable keeps the manure solid and moist, and checks the heating. This, however, is not a sanitary way, especially for dairy cows. But in some cases, for instance, where calves, colts, sheep, or steers are fed loose in a stable, the manure might be allowed to collect the entire winter. As soon as the animals are taken out, the manure should be hauled to the field and spread. In general, *it is better to clean the stables out daily and haul the manure directly to the field.* This is the very best practice wherever possible, and saves the manure more perfectly than any other method. Sometimes the ground is too wet or the fields, too hilly, or other causes arise which prevent hauling, and make it *necessary to store the manure.* In such cases, a good compact pile should be built where water can be applied to keep it from getting hot. The pile should be made solid and so deep that the hardest rain will not soak through. This can be accomplished by making the piles from five to six feet high. A basin-like depression with the bottom and sides of cement is a good place to store manure. A roof is not so necessary. *When manure is exposed in loose shallow piles, the losses of plant food in a few months may be at least one-half of the total in the manure.*

It has already been said that *manure should be hauled to the field as soon as possible. It should also be spread at once.* It is very wasteful of both fertilizer and labor to dump

manure in piles and let it lie there for weeks before spreading it. Every rain will wash out some soluble material into the soil just around the pile. This will give a **spotted meadow or grain field** the next season. The plants growing where the piles stood will grow more luxuriantly and mature later than the rest of the field. Spreading with the **manure spreader** is a good way to distribute the manure. On heavy soils

FIG. 82. — A manure spreader in action.

containing much clay more benefit will be derived from **fresh manure** than from those that are well rotted. Fresh manure warms these cold soils and makes them more porous. On light sandy soils, on the other hand, **manures that are well rotted** will be most beneficial. Fresh manure has a forcing effect and tends to produce stems and leaves at the expense of fruit and grain. It is, therefore, better for early garden truck, grasses, and foliage plants than for cereals. *The latter may lodge easily if grown directly after manuring.* Corn is usually benefited by liberal applications of fresh manure. In fact, when in doubt as to where to apply the manure, " use it on corn." *It is better to apply small quantities of manure to the land often rather than large quantities not so often.* Ten to fifteen tons per acre, once in three years, of manure of

average composition will replace the plant food removed by the average crops. This is not considered a heavy application. *With a manure spreader a small quantity of manure can be made to cover a large area of ground,* and thus, if desired, the fields can be treated oftener with smaller amounts.

By **green manuring** is meant the plowing under of green plants so that in their decay in the soil they will add humus and produce carbon dioxide. The addition of humus aids the soil in retaining moisture, and at the same time puts it in better physical condition for cultivation; thus sandy soils retain water better and clays become warmer and more workable. Any kind of plant can be used for a green manuring crop. *It is better to plow under weeds while they are green than to let them go to seed.* Rye is a common green manuring crop. It is sown in the autumn and plowed under in the spring. *The best green manuring plants are the legumes, because they add nitrogen, and this is what most soils need.* Clovers, soy beans, vetches, and cowpeas are all good. Red clover is the most common one used. These plants produce good results even when the crop is harvested, and only the stubble turned under. But because of the large root systems of alfalfa and red clover, their stubble leaves much more nitrogen than the roots of the pea or bean. **No plant but the legume adds nitrogen to the soil.** *All others can give back to the soil only the nitrogen which they obtained from it in the first place.* Usually it is better, in systems of animal husbandry farming, to cut the clover for hay and return the manure to the land. In this way the farmer gets the benefit of the feed and nearly all the plant food goes back to the land, *provided he saves the manure.*

QUESTIONS

1. What is the value of the manure annually produced in your state and how does it compare with the value of the dairy products?

2. What is a cold manure? What is its approximate water content?

3. How many pounds of nitrogen and phosphorus are there respectively in a ton of manure?

4. What factors influence the composition of manure?

5. Name two methods of increasing the value of manure.

6. How is the value of manure lost?

7. How would you prevent harmful fermentation?

8. How would you prevent leaching?

9. What is the best way to apply manure, and how many tons per acre should be applied annually?

10. What is meant by green manuring, and of what benefit is it?

CHAPTER XVII

PLANT LIFE AND WHAT IT PRODUCES

THE farmer rears plants and animals. He grows the plants on his land and his animals feed on the plants. He sells a part of his produce to those who need it and thus secures the means to buy clothing and tools for farming purposes, and to provide for the education of his children. *He begins his work with the seeds of plants.*

All living things possess certain properties in common. Careful study has shown that *the living substance in plants and animals is practically the same.* This living substance is called **protoplasm.** It is a clear, granular substance of about the consistency of the white of egg. It is this that grows, builds the plant and the animal, and performs all the functions of life. Exactly what it is, is not known, but it is closely allied to some of the complex bodies which the plant can build, as, for example, the proteins. Protoplasm is found in every plant and animal, not in one large mass, but in thousands of little parts called **cells.** These are so small that the compound microscope must be used in order to see them. The walls of different cells vary in thickness. They separate the cells from one another and enable them to maintain their form. In most of the common animals some of the cells have thick walls that are filled with mineral matter. These make up the skeleton, while the rest of the cells are soft-walled. *In plants most of the cells have fairly firm walls.* As long as the protoplasm is alive it is at work. One of its peculiar functions is to manufacture new protoplasm out of various sub-

stances which are not of themselves alive. The new protoplasm thus formed is alive and like that which produced it. *A plant or animal which is increasing in size and weight is making new protoplasm.* Among the other activities of protoplasm are the absorption and assimilation of food and the storing of reserve foodstuffs in the seed for future use, also the building of new cell walls, the secreting of substances for protection and the repair of injuries that may have been received, and the elimination of various waste products that are constantly being formed.

A knowledge of the seed and **seed germination** is of fundamental importance. On examining a kernel of corn it will

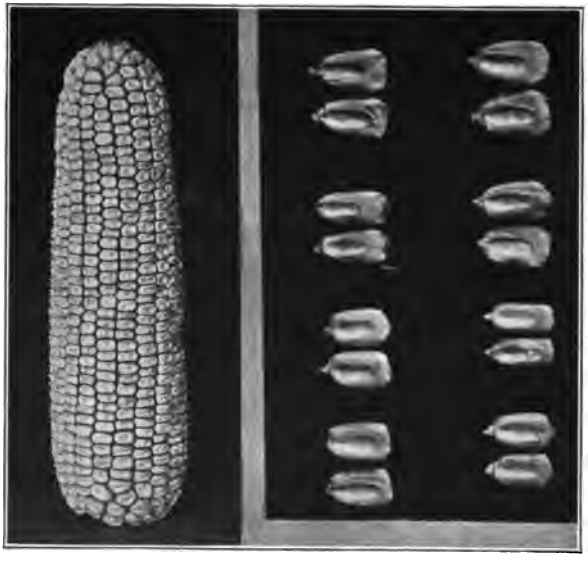


FIG. 83. — A good ear of corn. The kernels are of good shape.

be seen that it contains *a store of food material that is intended to nourish the young plant till it has sufficiently developed so that it can draw sustenance from the soil and the air.* It is to

be borne in mind that *plants live to produce seed,* which will in turn reproduce their kind. If a kernel of corn is soaked in water overnight, it can easily be separated into its distinct parts. There is the tough outer coating of the seed; at the tip is the cream-colored germ; packed around the germ is a layer of food material, mainly starch; and outside of this there is a still harder layer made up of starch and protein. *The germ is the only part of the seed that sprouts.* It may, indeed, be called the baby plant. It is the living protoplasm of the seed. The other materials are placed near the germ for the sole purpose of supplying it with food when it first begins to grow. *Every seed that is to germinate must have this living germ, this embryonic plant, in it; but the seed must also be supplied with* **moisture, warmth,** *and* **air.** Dry seeds do not germinate. In fact, *the way to prevent seeds from growing is to keep them dry.* This is well known from common experience. On the other hand, *wet seeds will either grow or rot.* They will grow if the germ is in good condition and warmth and air are furnished. Who has not seen the seed start in the shock or stack during a wet season, and who has not noticed that the beans rot when planted too early in moist, cold ground?

After the seed has been planted in sufficiently warm, moist ground, it absorbs moisture and swells. The seed coat is broken open, and the germ sends a root which starts downward into the soil. A little later a tiny shoot is put forth, which grows upward to form the stem, that is, the part of the plant above the ground. During this early period of growth the germ was fed by the starch and protein that surrounded it in the seed. That *the ground must be sufficiently warm in order that seeds may sprout and grow* is a matter of common experience. So oats sown during cold weather have sometimes remained unsprouted for a month, while

when sown in warm weather they came up in about one-third of that time. *Seeds of different plants require different temperatures to awake them and make them sprout,* that is, germinate. Thus seeds of wheat, oats, rye, and barley germinate when the ground is still quite cool, and so the

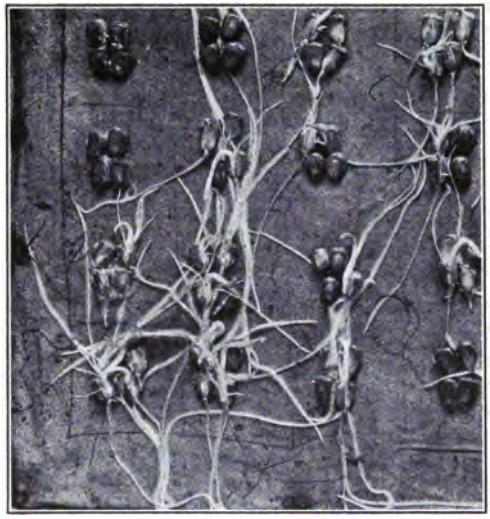


FIG. 84. — Germinating corn kernels.

farmer sows them in early spring. Corn requires warmer ground than wheat; and for seeds of cotton, cowpeas, or cucumbers the ground must be still warmer. A good farmer never plants these seeds until the soil has warmed up.

It has already been stated in Chapter XIV that the *air must have access to the roots of growing plants. Sprouting seeds also need air.* So if corn is planted in a field which is then overflowed with water, it will not come up. There are, indeed, some seeds that will germinate under water, like the seeds of rice and water lilies; but even in such cases, there must be oxygen dissolved in the water. *Packing the soil too*

closely around the seed will keep out the air and so hinder or prevent germination. This is particularly true of heavy clay soils. On the other hand, *packing loamy or sandy soils around seeds aids germination* by bringing the seeds in contact with water films surrounding the soil grains. These soils, being more porous, still permit sufficient access of air to the seed. *Some plants experience more difficulties than others in their efforts to come up,* that is, to reach the surface of the ground. For example, instead of sending up slender stems like the peas, or thin blades like the barley, they raise the entire swollen seed out of the soil, and try to lift a portion of the ground up with it. The bean is a plant of this kind. *In planting seeds these two types of behavior must be kept in mind.* Thus wheat, barley, peas, and corn belong to the first type; and they may therefore always be planted a little deeper than beans, radishes, beets, carrots, melons, clover, buckwheat, and squash, which are of the second type. The latter should never be covered more than five times their thickness. Other seeds may be covered somewhat deeper, but *no seed should be planted deeper than necessary to insure a proper supply of soil moisture.*

After the tiny root has started to grow from the seed, it becomes clothed with a downy fringe that looks like the finest silk. These delicate fibers are called **root hairs.** They absorb water and nutrients from the soil for the further growth of the plant. *During the life of a plant the root and stem that started in the seed normally continue to grow by division of certain groups of cells near their tips.* After a young plant has used up the material of the seed, it must gather its own **food supply.** A part of this, namely, *the carbon, is taken from the air by means of the* **leaf,** *but water and all of the other things necessary for growth are taken from the soil through the* **root,** which is consequently a very impor-

tant organ. It not only holds the plant firmly in position, but it also performs the very important functions of absorbing water, nitrates, phosphates, and various other necessary mineral nutrients. *All of these must be in solution in the soil water in order that they may be absorbed, for* **no solid matter can enter a plant.** The living membrane that lines the inside of each cell wall of a root hair will not let the finest solid particles pass, not even if they are finer than flour. *No part of the soil can act as plant food unless it is first dissolved.* Just as sugar and salt dissolve in water, so certain substances like the nitrates, phosphates, silicates, sulphates, chlorides, carbonates, etc., in the soil pass into solution in the soil water. In ordinary soil *the amount of material dissolved in the soil water is quite slight,* that is to say, soil waters are generally quite dilute solutions. Indeed, it would take several thousand pounds of soil water to carry to the plant one pound of calcium, phosphorus, or any other single element

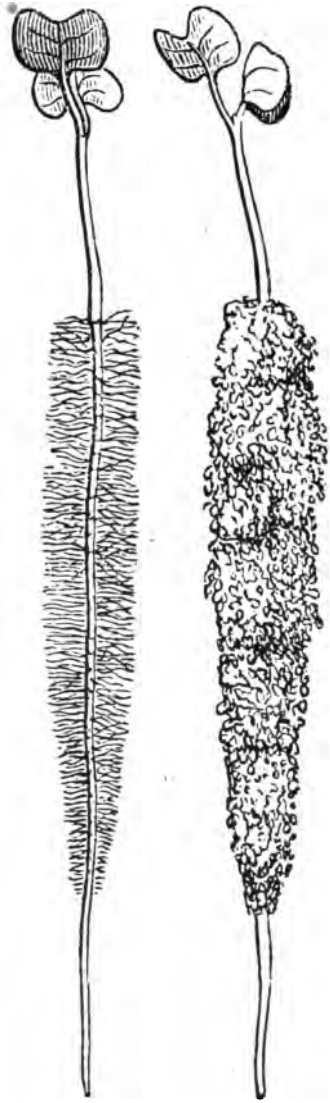


FIG. 85. — Root hairs. On the right the soil has not been washed away.

essential for vegetable life. *It is this* **soil water,** *then,* **with its plant food in solution** *that is absorbed by the root hairs on the roots and thus gains its way into the plant.* To accomplish their work, the root hairs penetrate between and wind around the soil particles, lying in the film of water that surrounds the solid grains of the soil. Moreover, as already mentioned in Chapter XIV, these rootlets, as they grow, continually give off carbon dioxide. This is absorbed by the soil water, which thus becomes a better solvent for many of the ingredients of the soil.

The **sap passes upward** from the root to the leaf *as a result of the action of capillarity and evaporation of water from the surface of the leaf.* The height to which sap may thus travel is sometimes very considerable. In the case of the giant trees of California, for instance, the distance from the root hairs to the leaves amounts to hundreds of feet, and the work done in transporting the liquid to this height is great. It is interesting to observe the upward passage of the liquid in the stem of a plant. This may be done by placing the stem of a white snowball or a lily in a vase of red ink. The red solution passes up through the stem and colors the flower. The **stem of a plant** is a system of tubes formed by continuously connected cells, some of which carry simple compounds in solution to the leaf, while others carry food, made in the leaf, back to the root.

It is fortunate for the farmer and for the food and clothing supply of all mankind that **the plant derives more material for its solid substance from the air than from the soil.** In every hundred pounds of dry plant material there are usually less than three pounds that have come from the soil. The grains of corn, wheat, rye, and rice consist chiefly of *starch.* Other plants, such as the sugar beet, are rich in *sugar,* while seeds like flax and cotton contain much *oil* and *fat.* Starch,

sugar, oil, and many other substances in plants are made in the leaves from the carbon dioxide taken from the air, and the water, nitrogen, phosphorus, potassium, and other elements which are absorbed from the soil in the form of salt solutions by the roots, as already explained. The carbon dioxide of the air enters the leaf through very small openings, most of which are located on the under side of the leaf. It then passes into the cells, where it comes into contact with **chlorophyll**, the substance to which the leaf owes its green color. Here a wonderful chemical change takes place, for *when the sun is shining on the leaf the carbon dioxide and some of the water in the cells act on each other, forming* **starch and oxygen.** The latter gas is exhaled by the leaf. The chemical equation expressing this change has already been given in Chapter IX, which see. Thus it is that *carbon is taken from the air* and retained by the plant. The starch and sugars formed in the leaf are plant food. *All parts of the plant, including the stem, branches, leaves, and roots, are nourished by food formed in this leaf laboratory from the elements gathered from the air, the soil, and water.* The formation of chlorophyll and *the conversion of carbon dioxide and water into oxygen, starch, and sugar cannot take place without sunlight.* The plant is like a machine run by a motor, the sun. All parts of the machine may be perfect and in readiness, but work will not proceed unless the motor is in action.

It has previously been stated that water evaporates from the surface of the leaf. It requires a great deal more water to carry food from the roots to the leaf than finally remains as an integral part of the tissues of the plant. This surplus water is transpired or given off by **evaporation from the leaves**, and is thus returned to the air. *Plants use several hundred pounds of water in making a single pound of dry solid matter.* Most plants thrive best when they have much

sunshine and frequent showers. In wet seasons they may suffer from want of sunlight, and in dry seasons from lack of water. The food prepared in the leaf goes to all parts of the plant in the form of sap. It passes through the soft fibrous layer called the cambrium, which is between the hard outer bark of the tree and the woody trunk. It is diffused throughout the plant and causes the cells that compose it to increase in size and number; that is, it causes the plant to grow. This sap may be obtained in large quantities from some trees, as, for instance, from the sugar maple, which is the source of maple sugar and maple sirup.

The leaf, which is the world's food laboratory, gives off oxygen in the sunlight as it synthesizes starch from carbon dioxide and water. But *in the absence of sunlight* **the leaf** *exhales carbon dioxide. It really requires oxygen to live, and breathes in that element just as animals do.* But during the time that the light falls upon the leaf, the carbon dioxide that would be exhaled is utilized instead, in making starch and sugar. Consequently only in the absence of sunlight, that is, when the workshop of the leaf is closed, does one really become aware of the fact that *the plant breathes in and uses oxygen, and gives off carbon dioxide* just as an animal does.

The plant is the source of food for all animals, including man. Man usually eats some meat, but this is derived from an animal which either received its nourishment from other flesh or from vegetable materials. In all cases, **the ultimate source of food is the plant.** The food materials produced by plants may be divided into two great classes: (1) the *non-nitrogenous,* and (2) the *nitrogenous.* The first group includes the carbohydrates and fats, while the second group, as the name implies, consists of nitrogen-containing bodies of which the proteins are by far the most important.

The **carbohydrates**, as already stated in Chapter IX, *include the starches, sugars, gums, pectins, celluloses, and closely related bodies.* These form by far the largest proportion of the organic compounds in the plant. In many plants and food materials the carbohydrates are present largely as starch. *About 75 to 80 per cent of the dry matter stored in a farmer's haymow and grain bins is made up of carbohydrates.* The activities of plant life are largely devoted to the production of carbohydrates; and their consumption by animals is correspondingly extensive. It should be remembered that these compounds are made of oxygen, hydrogen, and carbon, which are derived wholly from the air and water. The chemist in his analysis of plant tissues always separates the cellulose from the rest of the carbohydrates, as being of little value to animals. This crude cellulose he calls **crude fiber**, and to the rest of the carbohydrates he gives the name **nitrogen-free extract.**

FIG. 86. — Potato starch grains.

Starch is a widely distributed and abundant constituent of vegetable tissue. It is found in all the green parts of leaves, and accumulates in those organs of plants which serve as stores of reserve material. It occurs especially in roots, fruits, and seeds. In some seeds as much as 60 to 70 per cent is present. *Probably only water and cellulose are more abundant than starch in the vegetable world.* Starch does not exist in solution in the sap, but it is found

in the interior of plant cells in the form of minute grains, which have a shape, size, and structure characteristic of the seed in which they are found (see Figs. 86, 87, and 88). Potato starch grains are large, being about 0.003 of an inch in diameter, while those of wheat are smaller, about 0.001 of an inch in diameter, and resemble in outline a thick burning glass. *These differences in the size and shape of starch grains furnish the most important means of detecting adulterations of one grain with another,* as, for example, when corn flour is mixed with wheat flour. Unless changed by some chemical means, *starch is not soluble in water.* Prolonged treatment with hot water causes the starch grains to swell and burst, thus forming **starch paste**, which is used as an adhesive. Starch is spoken of as potato starch, wheat starch, rice starch, etc., according to the name of the plant from which it has been obtained. **Potato starch** is prepared by rasping the potato with water and passing the product through fine sieves. The coarse particles remain on the sieve, while the starch passes through the meshes in suspension in water. The turbid liquid is allowed to stand in vats till the starch has settled to the bottom. *All kinds of starch have the common property of being colored blue when treated with iodine.* Starch is used as human food, and it is also extensively employed in laundries. Applied as a paste to linen, it causes

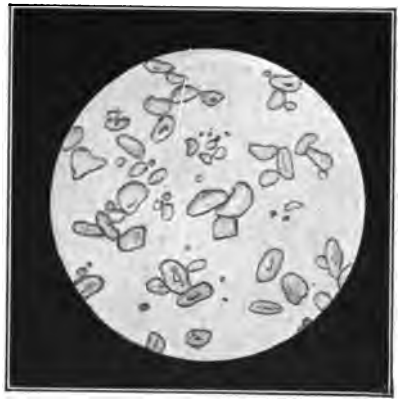


FIG. 87. — Pea starch grains.

the threads to adhere together, thus making the fabric stiffer. The hot iron of the laundry converts some of the starch to dextrins, which impart a gloss to the linen or cotton fabrics. **Dextrins** have a sweet taste, and they are the same substances that are formed from starch when the loaf of bread is baked. In the crust and in toast they are particularly abundant.

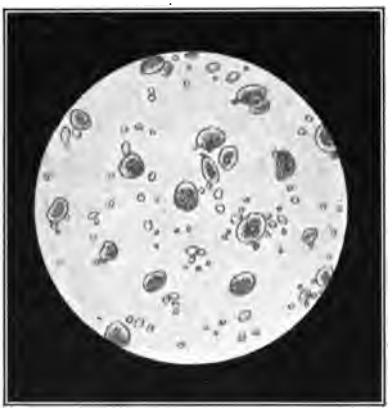


FIG. 88. — Wheat starch grains.

The **sugars** are very valuable carbohydrates, although they occur in plants in less amounts than the starches. They are found only in small quantities in hays, while in seeds their amounts are scarcely appreciable. In fruits sugars are more abundant. They are also found in solution in the sap of young plants, hence the sweet taste of this juice. *The most important sugar, commercially considered, is sucrose, also known as* **cane sugar** *or* **beet sugar** (see Chapter IX). As a human food it is widely used, and its manufacture and sale constitute a prominent industry. In some plants sucrose occurs in abundance, as in the sugar cane, which grows in the south. The sugar beet which thrives in northern climes is to-day a very important source of sugar. Sucrose also exists in sorghum and in considerable quantities in ordinary field corn, sweet potatoes, and pineapples. Fruits generally contain an appreciable amount of this sugar mixed with other sugars and organic acids. As already pointed out in Chapter IX, *dextrose, also known as* **glucose** *and* **grape sugar,**

is a very important article of commerce, usually appearing on the market in the form of certain candies and sirups, though it is also present to some extent in molasses, which see. Corn sirup is glucose prepared by the usual method, *i.e.* by heating starch with dilute sulphuric acid; see Chapter IX, where the behavior of dextrose and other sugars toward Fehling's solution is also described. **Fructose** occurs in fruits, and **maltose** is formed from starch when seeds germinate. These sugars too have been described in Chapter IX.

The **pectins** are closely related to the carbohydrates. They are the cause of the jellying of fruits such as apples and currants. They exist in greater abundance in fruit that is not quite ripe, and consequently this is commonly selected for **jelly making** in preference to ripe fruit. When such unripe fruits are cooked, the pectins take up water and are changed into gelatinous bodies which set to a firm mass on cooling.

The tough or woody parts of plant tissues consist of **cellulose.** *Wood, cotton, linen, hemp, flax, etc., when freed from mineral matter, are almost pure cellulose,* as already stated in Chapter IX. *Cellulose is insoluble in water,* and is not readily attacked even by strong acids, as previously explained, see Chapter IX. *In tables of the results of food analysis, the term crude fiber is commonly used. This material is largely cellulose.*

Plant tissues also contain **fats and oils,** and what has been said in Chapter IX as to the chemical nature of such compounds should be carefully reviewed in this connection. It will be recalled that though fats and oils, like the carbohydrates, consist of carbon, hydrogen, and oxygen, yet they are an entirely different class of compounds. The *fats and oils are mixtures of esters of oleic, stearic, palmitic, and other fatty acids with glycerine acting as the base.* The glycerine is set

free when the fats are heated with caustic alkalies, as in the process of soap making. Fats and oils contain a very much higher percentage of carbon than the carbohydrates. The organic fats and oils met in common life are partly of animal and partly of vegetable origin. So *tallow, lard, butter, goose and chicken grease are* **of animal origin**; while *olive oil, cotton-seed oil, linseed oil, etc.,* **are obtained from plants.** *Fats and oils are readily soluble in either chloroform, carbon tetrachloride, carbon bisulphide, gasoline, or ether.* This fact is utilized in *removing grease spots from clothes,* but also in *determining the fat content of various products* by means of chemical analysis. When a ground food material of either animal or vegetable origin is leached with ether, gasoline, or chloroform, the fats and oils it contains are dissolved. When afterward the clear solutions thus obtained are carefully evaporated to dryness, the residue which remains consists of the fats and oils plus such minor amounts of other substances as are slightly soluble in the solvent employed. When seeds are ground and then extracted with ether, the material dissolved in the latter consists of almost pure fats and oils; but when hays, straws, and green plant tissues are thus treated with ether, the extract contains the fats and oils plus other substances, like chlorophyll, which are also soluble in ether. *In tables showing the composition of foods there is commonly a column headed* **ether extract**; *the figures in this column represent practically pure fat in the case of seeds, but in the case of hay or straw they will represent fat plus minor impurities that have also been extracted along with the fat.* The amounts of fats and oils found in plant tissues vary greatly with the nature of the plant and also with the parts of any one particular plant. In general, *the seed contains much more fat than the leaf, stem, or root. Fleshy roots like potatoes and beets, for example, contain but little fat.* The common farm seeds, such

as corn, oats, barley, rye, and wheat contain from two to five per cent fat; while flaxseed (linseed), cotton seed, and peanuts run as high as thirty to forty-five per cent. The oils from some of the seeds are of great commercial value. Thus *cottonseed and olive oils are important articles of food, while linseed oil* (compare Chapter XII) *is used in enormous quantities in the paint industries.*

Some plants produce **acids** *in abundance.* In Chapter V attention has already been called to this fact. Thus *citric acid* occurs in lemons and other citrous fruits; *malic acid* is present in sour apples, currants, mountain ash berries, etc.; *tartaric acid* gives the tart taste to grapes; and *oxalic acid* may be found in rhubarb, oxalis, and the jack-in-the-pulpit. *The ordinary farm seeds, roots, and fodders, however, are free from acids; and it is only when they are subjected to fermentation that acids are formed from them.* A familiar example of this is the production of **silage.** Here fermentation changes the sugars of the fresh corn mainly to acetic and lactic acids (see Chapter IX for a description of these), and it is the volatile acids that give to silage its sour smell. The sour smell of the ill-kept kitchen garbage pail is also due to volatile organic acids produced by fermentation. What has been said about organic acids in Chapter IX should be carefully reviewed here.

The **proteins,** which are very complex compounds, are absolutely necessary to the life of every cell. They consist of carbon, hydrogen, oxygen, nitrogen, sulphur, and sometimes phosphorus about in the proportions indicated in Chapter IX. *It is for the purpose of building up these proteins that the root of the plant must have access to nitrates, sulphates, and phosphates. If any one of these is lacking, the plant ceases to grow.* Just as there are various carbohydrates, so there are several kinds of proteins. A familiar one is the

albumen of the egg. This will dissolve in water and coagulate when heated. The process of the **coagulation of albumen** is really a remarkable change which is by no means fully understood. Some very common matters of observation really depend upon the chemical, physical, and physiological properties of the proteins. So, for example, the farmer's boy secures a tenacious cud of gum from fresh wheat kernels, the housewife watches the strings of coagulated albumen separate from the cold water extracts of fresh lean beef when the liquid is boiled, or observes the white mass harden as the egg is dropped into boiling water. *The seeds of plants are richer in proteins than the stems.* Wheat grain contains from eleven to twelve per cent of protein, while in the straw there are only from three to four per cent. Some seeds are particularly rich in proteins. So, for instance, *peas and beans may contain from sixteen to twenty per cent of protein.* The proteins that are soluble in water are called albumins. They are found in the juice of plants, but also in blood, milk, and especially in eggs. The **globulins** are another class of proteins which *occur most abundantly in seeds.* While they are insoluble in water, they may be dissolved by treating the ground seed with a five per cent solution of common salt. *There are numerous other proteins* which cannot be mentioned here. *When heated with soda lime all proteins give off ammonia, which fact is frequently used to detect their presence.* Certain color reactions to which attention will be called in connection with the laboratory experiments (Chapter XXII) also serve for this purpose.

In young growing plant tissue there is found a group of nitrogenous substances called "*amides.*" In complexity of chemical composition, these lie between the nitrates and the proteins. It is in the form of these amides that the nitrogen of the growing plant is carried around to the various cells.

Nitrogenous compounds of this kind are far more abundant in hays and roots than in seeds. A considerable part of the nitrogen contained in plants cut for hays, like timothy, the clovers, and alfalfa, is in this form.

QUESTIONS

1. What name is given to the substance common to all living matter? In what is it contained?

2. What name is given to the living part of the seed? What three things are necessary to make a seed sprout?

3. What effect does temperature have on the germination of seeds?

4. How do seeds differ in the way they come through the ground?

5. What is the action of the root, and how does food material enter it?

6. What proportion of the solid substance of a plant comes from the air?

7. What does the leaf of a plant do? What substance in the leaf is especially active in sunlight?

8. What is the function of sunlight in making starch? Compare the breathing of a plant with that of an animal.

9. Name three substances that are carbohydrates. How would you test for starch?

10. How does cane sugar differ from starch? What substance in fruits causes their jellying? How does cellulose differ from starch in its properties?

11. What elements are contained in fats? What dissolves fats? What does the ether extract in the process of food analysis contain?

12. What elements are found in protein? Why is nitrogen such an important element in the soil? Name a protein.

13. What two acids are found in silage?

CHAPTER XVIII

THE ANIMAL AND ITS FEED

ANIMALS cannot feed on the mineral matter in the soil nor the carbon dioxide in the air. Plants live on both. *The chief usefulness of plants to man lies in their ability to convert the minerals of the soil and the carbon dioxide of the air into substances that will nourish him and his servants — the domestic animals.* Plants form the natural food of the animal of the farm.

The bodies of animals consist of flesh, fat, bones, blood, teeth, hair, etc.; nevertheless, the chemist would say that *the same classes of substances are present in the bodies of animals as are found in plants;* namely, water, ash or mineral matter, proteins, fats, and carbohydrates. Each of these substances from the plant may be utilized by the animal in forming somewhat similar materials in its body. But plants contain important substances, like cellulose and the starches, which are not present *in animals. The starches represent one of the most important foods of animals.* Instead of adding starch to its own body unchanged, the animal converts it into animal fat, or uses it as a fuel to produce animal heat. *Cellulose, being a very resistant material, is utilized by animals as food only to a slight extent.* Most of it that is eaten passes out of the intestine as a part of the feces. *More than half of the weight of the body of an animal consists of water; and besides the water which animals drink, they eat food which is in a large measure composed of water.*

The dry **mineral matter** which remains when the bodies of plants or animals are burned is called the **ash**. The nature of plant ash has already received attention. *The ash of an animal comes mainly from its skeleton.* While the skeleton usually represents only about 8 per cent of the total weight of an animal, nevertheless 95 per cent of all of the ash of the body is in this earthy framework, which consists principally of calcium phosphate. *The ash of the animal always contains all of the mineral elements that are necessary for the growth of the plant.* The same mineral constituents that a plant must have in order to be able to develop are also required by the animal in its process of growth. Animals generally contain more sodium and chlorine than plants, for *the blood is rich in common salt. There is generally an abundance of ash in the common feeds to supply all of the mineral matter the animal needs.* But *pigs fed on grain alone*, without access to a pasture, *may to advantage receive wood ashes* in their feed, for this will furnish them more lime, and thus aid in bone formation and other ways. *Laying hens having no opportunity to range need extra mineral matter* in the form of oyster shells, bone, or limestone. These should be suitably ground. They are used by the hen in making the shell of the egg.

The animal is essentially a protein structure, while the plant is a carbohydrate structure; for the animal is rich in proteins, and the plant usually does not contain so much of these compounds. The **proteins** of the plant are used by the animal to make lean meat, muscle, blood, nerves, brain, hoofs, hair, and the casein of milk. It is also very probable that fat, which is non-nitrogenous in character, may be formed from proteins; but this question is still disputed. Proteins serve as a source of energy. In the case of a dog that eats only meat, probably the greater part of the protein eaten is not stored, but used as fuel for producing body heat. *An*

animal can get along without carbohydrates or fat, but it must have a certain amount of protein in its daily ration, for the cells of the body are constantly losing nitrogen, which can only be replaced by protein.

The **fats**, it will be recalled, are free from nitrogen. Those consumed with the food may be deposited in the tissues of the body without essential change. Small deposits of fat, indeed, occur in every organ and cell, but the amount held in reserve varies greatly with the nutritive condition of the animal. So a lazy, overfed pig will have great reserves of fat, while a hard-worked horse will not show such reserves. The fat of the food is either burned in the body to furnish heat and mechanical energy, or it is stored as reserve material.

The **carbohydrates**, especially *starch, sugars, and cellulose, form by far the largest part of all food of vegetable origin.* They are not stored permanently in the body, but as they are oxidized in the system they furnish heat or mechanical work. There is always a small amount of sugar in the blood, — less than one per cent, — but after a meal the liver and muscles contain an abundance of the carbohydrate **glycogen**, *which is really animal starch.* But this is only a small portion of the animal's daily intake of carbohydrates. When consumed in excess of the immediate requirements, *carbohydrates are changed to fats and stored as such.* As a practical illustration of this truth may be mentioned the well-known value of corn meal, which is particularly rich in starch, as a fattening food.

The animal body performs work. Even an idle horse expends some energy in the circulation of the blood and the digestion of food. The more work the horse does, the greater the amount of energy he must expend; and so the horse uses a part of the energy represented in the food he eats for the production of motion, just as a steam engine uses coal in furnishing power. In addition, *a certain amount of food is*

oxidized in the animal body daily to maintain its temperature. The temperature of domestic animals and man is about 38° C., which is much above the prevailing temperature of any climate, and to *maintain this temperature fuel is required.* All of the foodstuffs — carbohydrates, fats, and proteins — may serve as fuel in the animal body to produce heat and other forms of energy; but when fats and carbohydrates are available, the proteins of the tissues are not normally burned for this purpose. Only when carbohydrates and fats are lacking will the animal burn up its proteins; and so, because the latter are relatively expensive, *it is more economical to have a plentiful supply of fats and carbohydrates in the diet as sources of heat and other energy, providing sufficient protein, however, when new tissues are to be built,* as in growth, or when milk is to be produced, as in the case of the milch cow. A working horse needs some protein, but uses mainly carbohydrates and plant fats for the production of heat and work. *A pound of fat will produce 2.4 times as much heat as a pound of carbohydrates;* which will not seem strange when it is remembered that fats contain relatively much more carbon than is present in carbohydrates. A pound of protein will produce about the same amount of heat as a pound of carbohydrates.

The important process by which the food of animals or man is rendered capable of being absorbed into the system and used in building up or renewing the tissues of the body is called **digestion.** So, for instance, corn meal or hay cannot directly be transferred to the blood. These must first be brought into solution before they can pass out of the digestive tract into the blood and lymph; for *just as no solid particles can pass into a plant through its rootlets, so no solid particles can pass through the walls of the alimentary canal into the body.* Furthermore, *just as the rootlets of a plant*

give off carbon dioxide which aids the soil water in dissolving plant food, so the walls of the alimentary canal secrete certain substances which aid in liquefying food so that it may be absorbed. The process of getting food ready for absorption is partly a mechanical one, but it consists mainly of a series of chemical changes which are produced to a large extent by the action of bodies called *enzymes*. These are peculiar substances which are produced by all living plant or animal cells. They are in general non-crystalline, slimy, soluble substances containing the same essential elements as the proteins; but their exact composition is still unknown. **Enzymes**, *also called unorganized ferments, have the power to bring about certain complex chemical changes at ordinary temperatures.* For example, the enzyme in the saliva causes starch to change to the soluble sugar, maltose. If starch were to be changed to sugar in the laboratory, the chemist would have to add an acid and apply heat. Another peculiarity is that an enzyme that can act on starch has no effect on proteins or on fat. So the starch-splitting enzyme has no action on the casein of milk. *Enzymes are easily destroyed by heat.* The temperature of boiling water stops their action.

The process of digestion begins in the mouth. The first step is **mastication**, by which the food is subdivided and crushed by the action of the teeth and at the same time well mixed with saliva. The latter is a very dilute solution which turns litmus paper blue and contains the enzyme **ptyalin**, which has the power of converting the insoluble starch of the food to soluble sugar. This change begins in the mouth and continues in the stomach for a time. But when the acid juice of the stomach comes into contact with the ptyalin, this enzyme ceases its action. *Proteins and fats are not attacked by the juice of the mouth.* It is estimated that *an ox or a horse secretes one hundred pounds of saliva daily.* This

fluid also prepares the food for swallowing. After mastication the food passes down the gullet into the stomach. *In the case of man, the horse, and the pig, the stomach is a single sac, and true digestion begins here at once. In ruminants* (cud chewers), *like the ox and the sheep, the stomach consists of four sacs, and not until the fourth one is reached by the food does true digestion begin.* These sacs are used for storage, and they are rather complicated in structure. The large paunch of the ox, which is the first compartment reached by the food after it leaves the mouth, has a capacity of 50 to 60 gallons. It is in the paunch that the food is stored; and because of its coarse nature, it is returned to the mouth later for more thorough dividing and mixing with saliva. This is what is called "chewing the cud."

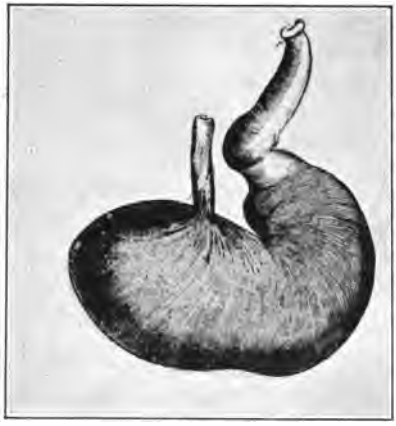


FIG. 89. — Stomach of a horse.

The action of the saliva on the food in the paunch is thus continued much longer than would be the case in the true stomach of man or the pig. When the food reaches the fourth stomach in the ox, or the first one in the pig, true digestion begins. The active *digestive fluid* comes from cells in the mucous membrane that lines the stomach. It is a watery fluid containing various salts, free hydrochloric acid, and the two enzymes, **pepsin** and **rennin**. *This combination of pepsin and acid is the effective agent in the digestion.* It acts on the proteins of the food, dissolving them

by converting them into simple nitrogenous bodies. This enzyme, like ptyalin, works best at the body temperature, 38° C. It is said that the stomach juice of the sheep has a low acidity, 0.1 per cent, while that of the dog is much higher. Chemically *the results accomplished in the stomachs of all farm animals are the same; that is, the proteins are changed to simple soluble nitrogenous substances.* The utilization of coarse feed by the horse is not as complete as by the ox, for the reason that with the former there is no preliminary remastication before the food material comes into contact with the gastric juice. *The purpose of the rennin in the stomach juice is to curdle the milk.* Here rennin acts exactly the same as when it is used in making cheese. As **milk**, the sole source of nutriment for the young, is a fluid which cannot be directly absorbed, it would not be so completely digested if it were not first converted to a solid mass, for it would pass through the intestine too fast. The rennin curdles it and gives the pepsin a longer time to act.

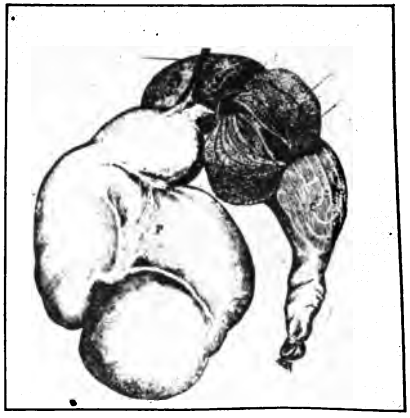


FIG. 90. — Stomach of a sheep.

When the food leaves the stomach, it enters the small intestine. At this point it is only partly digested. *The fats have not been changed,* and undoubtedly a considerable part of the proteins and carbohydrates is still to be acted upon. *The juice from the pancreas,* which is poured into the intestine at a point just below the junction of the stomach and small

intestine, *is the active agent in completing the solution of all the materials that can be digested.* It is a strongly alkaline solution and contains three enzymes, (1) **trypsin**, a protein-splitting enzyme, (2) **lipase**, a fat-splitting enzyme, and (3) **amylopsin**, a starch-digesting enzyme. By the action of this juice, the acid chyme from the stomach becomes alkaline, and the unattacked proteins, fats, and carbohydrates are reduced to simpler compounds, which are more soluble in the juice and capable of being absorbed through the intestinal wall. The fat undergoes but little change, but a part of it is converted to soap, the formation of which helps *to hold the rest of the fat in suspension as very finely divided globules that can be absorbed.*

The proteins, carbohydrates, and fats of the food are gradually digested by the processes just described. *Throughout the length of the long intestine absorption proceeds rapidly.* Water, salts, and the products of digestion pass from the intestine into the circulating lymph and blood. The *sugars* are carried to the liver, converted into glycogen, and stored temporarily, being doled out again to the blood in constant quantities to be carried to the remotest cells of the body for fuel. The *proteins* are carried to all the cells of the body, where they serve in the work of repair and new growth. The *fats* are either deposited in the various tissues, or are used as food by the hungry and moving muscle cells.

In the lungs the blood is supplied with oxygen. Here purple venous blood is changed to scarlet by the oxygen which it absorbs. At the same time, a considerable quantity of carbon dioxide, which was formed in the burning of fuel in the cells, is brought to the lungs by the blood stream and given up to the air. *Inspired air contains about 21 per cent of oxygen and 0.03 per cent of carbon dioxide,* while in expired air there is approximately 16.5 per cent of *oxygen and 4.4 per cent of*

carbon dioxide. Though oxygen is absorbed in the lungs, it is not here that the process of oxidation of the carbon and hydrogen content of the food takes place. *The blood acts simply as a carrier of oxygen and the combustion itself actually takes place in the remote tissues.*

The nutrients of foods are not wholly digested. A part passes through the animal without having been liquefied and dissolved by the digestive juices and thereby made available to the animal. Foods differ in this respect. Those rich in cellulose, such as *hays, straws, and the coarse parts of grains, will leave much greater indigested residues than whole grains or the interior parts of grains. Milk is completely digested.* Animals likewise differ in their ability to digest coarse feeds. *The cow and the horse will make better use of hay and straw than the pig.* This is mainly due to the large storage apparatus in the former animals. Because of these **differences in the digestibility of feeds**, the availability of their nutrients naturally differs. This availability of the nutrients of foods is called the **coefficient of digestibility**. *It is the amount of nutrients that the animal can secure from a given feed.* For example, the coefficient of digestibility of the dry matter of corn meal is 89.6, and that of timothy hay is 57.9. This means that in 100 pounds of dry matter from corn meal, 89.6 pounds are available, while in timothy hay but 57.9 pounds are available. These figures are for the cow. Larger reference books on the subject of feeds and feeding must be consulted for tables giving such coefficients of digestibility of foods for the various domestic animals.

A knowledge of the relative digestibility of the several nutrients contained in the feeding material determines *how an animal ought to be fed,* that is, it suggests how to make a **feeding standard**. Such knowledge has been secured by many experiments with animals. It has also been found

in practice that *the feed of an animal may be varied within fairly wide limits, provided the ratio of digestible proteins to all other digestible organic matter is kept within certain limits.* Protein has special functions and less than a certain quantity would limit growth or production. For this reason *the ratio of protein to the other nutrients must be held within definite limits.* It will be recalled that fat and starch are used for fuel, and that a pound of fat can produce 2.4 times as much heat as a pound of starch. *The* **nutritive ratio** *is always expressed as amount of protein to amount of starch, including in the starch the fat reduced to a starch basis.* The nutritive ratio consequently becomes:

$$\frac{\text{The weight of digestible protein}}{\text{The weight of digestible carbohydrate} + (\text{digestible fat} \times 2.4)}$$

For example, the nutritive ratio of corn meal is obtained as follows:

100 lb. contain 7.9 lb. digestible protein.
66.7 lb. digestible carbohydrates.
4.3 lb. digestible fat.

$$\frac{7.9}{66.7 + (4.3 \times 2.4)} = \frac{7.9}{66.7 + 9.32} = \frac{7.9}{76.02} = \frac{1}{9.6}$$

The nutritive ratio of corn meal is therefore 1 : 9.6. This means that for every pound of digestible protein in corn meal there are 9.6 pounds of digestible carbohydrates and fat equivalent thereto. A *wide ratio* is one in which there is a large amount of carbohydrates in proportion to the protein. Oat straw has a wide ratio, namely 1 : 33.7. The corn ratio is *medium*, while the ratio of oil meal, 1 : 1.7, is called *narrow*.

The **Wolff-Lehmann feeding standards** for farm animals are the oldest and most used. They were constructed after many scientific and practical experiments on domestic

animals had been made to determine the proper amounts of protein, fat, and carbohydrates for a ration. The results of this work are embodied in what are commonly called feeding standards. A ration conforming to such a standard, *i.e.* one containing the proper amounts of digestible protein, fat, and carbohydrates, is called **a balanced ration.** *Feeding standards are not hard and fast rules, because there are differences in feeds and animals; but they are good general guides.* Below are examples of standard rations for several animals. For complete tables, texts on feeds and feeding should be consulted.

WOLFF-LEHMANN STANDARD

FOR 1000 LB. LIVE WEIGHT, DAILY

	DRY SUB.	DIGESTIBLE			NUTRITIVE RATIO
		Prot.	Carb.	Fat	
	lb.	lb.	lb.	lb.	
Cow, milk yield 22 lb. .	29	2.5	13	0.5	1 : 5.7
Fattening steer, 1st period	30	2.5	15	0.5	1 : 6.5
Horse, medium work . .	24	2.0	11	0.6	1 : 6.2

This table indicates that a cow yielding 22 pounds of milk of average composition will need daily 29 pounds of dry matter, which shall contain 2.5 pounds of digestible protein, 13 pounds of carbohydrates, and 0.5 pound of fat, for every 1000 pounds of her live weight.

Practical tests have shown that *animals generally thrive better when the ration is properly balanced.* A balanced ration is more economical for the farmer as well as better for the animal. *There is generally waste in an unbalanced ration.* If, for instance, an animal is fed a ration which is too high in

carbohydrates and too low in protein, it will consume more carbohydrates than it needs in order to get the necessary protein. Sometimes, under special conditions, depending on the feed on hand and the market value, it is necessary or desirable to feed an unbalanced ration. This should, however, be guarded against as far as possible.

Young animals require less food than older ones, to produce the same increase in their weight. Thus *it is more profitable to fatten a pig which is less than a year old than one nearly two years old.* **Growing animals** *need an abundance of protein and ash* in their rations. Narrow rations suit them best. Milk is a narrow ration, and it is well adapted for young growing animals. *The food requirements of fattening animals are quite different from this.* When fattening, an animal is not storing up protein, consequently its demands for protein are much less than in the case of a growing animal. In fact, *carbohydrates and fats can be abundant in the* **rations of fattening animals**, and a nutritive ration of 1 : 10 has actually been found satisfactory. This is much wider than that recommended by the Wolff-Lehmann standard, but it is now considered as sound practice by the best authorities.

Working animals *may also advantageously be fed rations of fairly wide nutritive ratios.* Both heat and other energy can be produced by animals from carbohydrates and fats, as already stated. Since this is true, no large supply of protein above that required for maintenance need be supplied. *The farmer always produces an abundance of carbohydrates, and for this reason it is fortunate that working and fattening animals as well can be properly fed with the farm-grown materials,* without the purchase of protein concentrates, such as gluten feed, bran, or oil meal. Horses working on the sugar plantations of the Fiji Islands receive 15 pounds of molasses daily, and their feed has a nutritive ratio of 1 : 11.8.

It is said that they do well on this ration. **Milch cows** are machines producing an abundant secretion of a highly nitrogenous character, for *30 pounds of milk contain one pound of protein.* For this reason, it is clear that their ration should be richer in protein than that of a work horse or fattening steer. It is conceded that *cows need at least 0.6 pound of protein to repair the daily wear and tear of their cells. This, plus what goes into the milk, makes 1.6 pounds.* Practice has established that an additional supply of protein, over that required merely for maintenance and for the milk, is necessary. *For milch cows a nutritive ratio of 1 : 5.5 to 1 : 7.0 is the general practice* and gives the best results. Young pasture grass, as good a milk producer as we have, is even narrower than this. **All farm animals required to work, fatten, or produce milk must receive grain or concentrated food as a part of their ration.** Bulky, coarse food, such as the hays, use much energy in being moved along the intestinal tract and will not give the best results when used alone.

QUESTIONS

1. Of what is an animal composed? How much water is there in a 1000-pound horse?

2. Of what use is ash or mineral matter in food?

3. What does an animal do with the starch and sugar it eats? What with the protein? What with the fat?

4. What is meant by "digesting" the food, and what is the name given to the substances that do this work?

5. Describe digestion in the mouth and in the stomach.

6. What curdles the milk in the stomach? Is this the same substance used in cheese making?

7. What happens to the food in the intestine?

8. What is the purpose of respiration, and why should we live in well-ventilated rooms?

9. Do foods differ in digestibility? What foods leave large residues in the intestine?

10. What is meant by the nutritive ratio of a feed?

11. From the tables in a text on feeds and feeding, calculate the nutritive ratio of the oat grain and alfalfa hay; also of milk.

12. From the same tables, calculate a balanced ration for a dairy cow weighing 1000 pounds and giving 30 pounds of milk daily. The materials available are oats, corn, gluten feed, and alfalfa hay.

13. What differences may there be in the ration of a fattening pig, a working horse, and a growing animal?

CHAPTER XIX

HUMAN AND ANIMAL FOODS

THE foods consumed by human beings are of vegetable and animal origin. The vegetable kingdom furnishes all of the necessary constituents of food in cheaper form than the meat of animals and without an excess of protein. The *distinguishing characteristic of vegetable foods is their richness in carbohydrates, a feature which offers a striking contrast to animal flesh, which contains practically no carbohydrates.* The only common food of animal origin containing a fair proportion of carbohydrates is milk with its content of milk sugar. The *disadvantage of vegetable foods* is that they are relatively *bulky* and as a rule but *incompletely digested,* for their chief nutrient, starch, is contained in a vast number of cells the walls of which are composed of cellulose, which, as already pointed out, is not easily dissolved by the digestive juices. *It is strongly in favor of vegetable foods, however, that they contain all of the various essential food constituents, and that in many vegetables these are present in well-balanced proportion.* Bulkiness is a property that is common to all vegetable foods. It is due chiefly to the large amounts of water and carbohydrates which vegetable matter contains. The **green fresh vegetables** are especially bulky; in fact, their bulk is out of all proportion to the small amount of nutrients which they contain. Four-fifths of the weight of fruits, green peas, beans, cabbage, carrots, radishes, etc., is due to water. It is probably safe to conclude, however, that *the proteins.*

fats, and carbohydrates in vegetables are equal in nutritive value to the same substances derived from the animal kingdom.

Bread *is made from the grains of cereals.* It is the basis of human *nutrition*—the staff of life. **Cereals** also enter largely into the rations of our domestic animals. The importance of cereals rests on the fact that they and the products made from them furnish the chief food constituents — proteins, carbohydrates, fats, and mineral matter — cheaply, abundantly, and in an agreeable form. These qualities distinguish the cereals as foodstuffs of the first rank. Nevertheless, when used alone, they are not well suited for the production of growth, and should consequently be accompanied by foods richer in protein, like milk or eggs. In the case of domestic animals, **mill concentrates** like gluten feeds and oil meals will serve to furnish the additional protein required. The *carbohydrates predominate in some cereals, like rice and Indian corn* (maize), and these are consequently prominent as sources of energy. **Rice** *and* **wheat** *have a low fat content, while the* **oat** *and* **corn** *grain contain as high as 5 per cent of fat.* Though the proportion of protein in cereals varies, yet the protein is extremely useful, for it furnishes the materials necessary for the growth and repair of tissue. *Wheat is the most important cereal, but rye, rice, corn, oats, barley, and buckwheat are also of great value.* The sticky character of the *proteins of wheat* peculiarly adapts it to the making of light, porous bread. This same property is also possessed by the *proteins of rye,* but none of the other cereals will make a light, porous loaf of bread. **Barley** is one of the most ancient of foods. It was highly esteemed by the athletes of ancient Greece. The amount of water in cereals varies but slightly and averages about 11 per cent. The **general composition of the cereals** is about as follows: water, 11 per cent; protein, 10 to 12 per cent; carbohydrates,

65 to 70 per cent; fat, 0.5 to 8 per cent; ash, 2 per cent. Such averages are useful, but they should not be applied too rigidly to any particular sample of grain. Thus the proteins of wheat may run as high as 17 per cent, or as low as 8.5 per cent, depending upon the climate in which the wheat was grown; but these figures are extreme and unusual. Most of the cereals contain only from 2 to 3 per cent of cellulose; but grains having thick outer coats, like oats and buckwheat, have a cellulose content of 10 to 12 per cent.

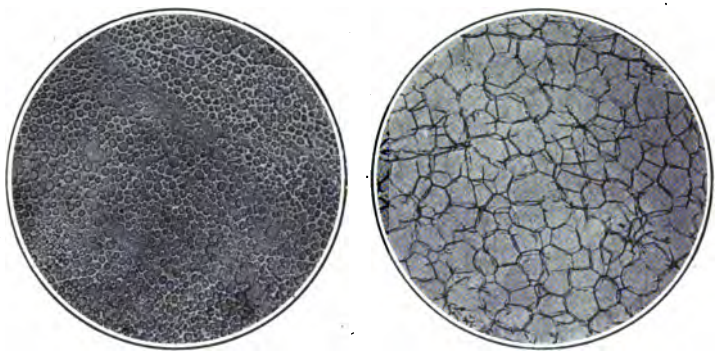
Before being used for human food, the cereals are usually ground to a powder. From this powder the coarser cellulose particles are sifted out, and the fine product so obtained is called **flour**, the term **meal** being applied to the coarser products. In thus sifting out the **bran**, *i.e.* the outer coating of the grain in the case of wheat, the quantity of mineral matter in the flour is greatly reduced. *For children it would be better if their bread were made from the coarser meal*, because the cellulose and mineral matter contained in it are laxative in their action. Twenty-five years ago the wheat was simply crushed and ground between millstones, and then crudely separated into three products by sifting. These three products were called *bran, flour*, and *middlings*. *Bran, which is so commonly fed to dairy cows to-day, is rich in mineral matter, especially in potassium and phosphorus, but it also contains about 15 per cent of protein.* All of these have been taken from the wheat kernel.

Foods are prepared for the table by **baking and cooking**, during which processes they are acted upon *physically*, *chemically*, and *bacteriologically*. *The chemical compounds of which foods are composed are altered to a greater or lesser degree by heat.* In boiling a potato, for instance, the cellulose of the cell walls unites to a certain extent with water; that is, it is partly hydrated, and so becomes softened. Further-

more, at the temperature of boiling water, starch is slightly changed to dextrin; but the change is only slight, because the temperature is not high enough to effect complete alteration. On the other hand when baked with a dry heat, more of the starch will be dextrinized, for the temperature of the oven is considerably above that of boiling water. The sugars are not altered appreciably by boiling water, but by dry heat they may be changed to a dark brown substance called **caramel**. During the formation of caramel, water is given off. *The dark color of the crust of the loaf of bread is due to caramel.* The fats are acted upon to some extent by heat. Some of the vegetable oils undergo slight oxidation, but these changes are not regarded as of any great influence on the food value of these substances. In general, the proteins tend to become less soluble by the action of heat, but not all proteins are thus altered. The white of the egg is coagulated by heat, but the main protein of corn meal is not. *Coagulated or hardened protein will be digested more slowly than raw protein.* But cooking produces such important physical changes, especially on starch and the cell wall, that its good effects more than counterbalance the unfavorable action on the proteins.

The physical changes produced by heat and water consist of a partial disintegration of the animal and vegetable tissues of foods. The cementing materials are loosened, and a softening of the tissues results. Often the action proceeds still farther in the case of vegetable foods, resulting in the disintegration of the individual starch granules. When foods are subjected to dry heat, their moisture is converted into steam, which causes the bursting of the cells. *Popping corn is a good illustration of this.* In baking bread or other cereal doughs, the starch grains are partially ruptured, and they are consequently more readily digested than when raw. Prolonged

and slow heat is most effective in improving the mechanical condition of the tissues. At too high a temperature, the natural flavoring materials in the foods are partly volatilized, and so, of course, lost. By boiling foods in water, salts are



(*A*) (*B*)

FIG. 91. — (*A*) Popcorn partly popped. (*B*) Popcorn fully popped. Highly magnified.

extracted from them, and so their palatability is lowered. This undesirable effect will be greatly lessened by adding salt to the water in proper quantity.

The bacterial organisms in foods are destroyed in the process of cooking, provided a temperature of 150° F. is reached and maintained for several minutes. The interior parts of foods rarely reach a temperature above 200° F. while cooking goes on. *Not only are bacteria destroyed by cooking or baking, but various parasites, such as trichina and tapeworms, are also destroyed.* But it should be remembered that cooked foods are easily reinoculated; in some cases more readly than fresh foods, because they are in a more disintegrated condition. In many instances, **bacteria and yeasts** *are of material assistance in the preparation of food, as in bread making, butter making, and the curing of cheese;* and it must

be borne in mind that in such cases the bacteria are apparently not at all harmful when eaten with the raw food.

In order to make flour available as food it must be cooked in a proper manner,. The simplest way is to mix it with water and bake it. This is **ship's biscuit**, a leaden, heavy mass. **Bread** *may be made light by having a gas, carbon dioxide, develop in the mass, which slowly pushes the particles apart and thus produces a light sponge.* The viscid nature of the gluten of the flour allows this to be done successfully. To produce this gas, yeast is commonly used. It consists of very minute plants, capable of breaking up sugar into carbon dioxide and alcohol. The carbon dioxide gas is the effective agent in lightening the dough. *Two-thirds of the bulk of the fermented bread is gas. Of the solid residue, about 40 per cent is water.* This is relatively much less than the water content of meat, which is from 70 to 80 per cent. Bread also contains a little more than 50 per cent of carbohydrates, about 7 to 8 per cent of proteins, and 2 to 3 per cent of fat and ash.

All **baking powders** *contain at least two ingredients,* (1) a carbonate, and (2) a compound which serves to liberate the carbon dioxide gas. The carbonate from which the gas is obtained is almost always **sodium bicarbonate**, $NaHCO_3$, commonly known as " baking soda " or " saleratus." The compounds that serve to decompose the baking soda are in general acidic in character, those in common use being **cream of tartar**, $KHC_4H_4O_6$, **tartaric acid**, $H_2C_4H_4O_6$, **calcium acid phosphate**, $CaH_4(PO_4)_2$, and **alum**, $NH_4Al(SO_4)_2$. These may be used separately or in combination. *All of these liberate carbon dioxide when mixed with baking soda and moistened, but the products left in the food differ widely in amount and character.* It is the nature and amount of the residue that remains in the bread which determines whether one baking powder is more desirable than another. In the **cream**

of tartar baking powder, the acid ingredient is potassium acid tartrate. This reacts with the sodium bicarbonate, liberating carbon dioxide and forming sodium potassium tartrate, which remains in the bread. In the **phosphate baking powder** the acid constituent is monocalcium phosphate and the residue left in the bread is a sodium calcium phosphate. The residues left in the bread or other foodstuffs from these two powders are perfectly harmless. In the **alum baking powders** the active agent which decomposes the baking soda is ammonium aluminum sulphate, although sometimes potassium or sodium aluminum sulphate is used. In any case, the aluminum hydrate formed in the reaction with the sodium bicarbonate is a harmful material which forms aluminum chloride with the hydrochloric acid in the stomach, and this salt hinders digestion in the same way that alum does:

$$\underset{\text{alum}}{2\,NH_4Al(SO_4)_2} + \underset{\text{"soda"}}{6\,NaHCO_3} = \underset{\substack{\text{aluminum}\\\text{hydrate}}}{2\,Al(OH)_3} + \underset{\substack{\text{sodium}\\\text{sulphate}}}{3\,Na_2SO_4} + \underset{\substack{\text{ammonium}\\\text{sulphate}}}{(NH_4)_2SO_4} + \underset{\substack{\text{carbon}\\\text{dioxide}}}{6\,CO_2}$$

In many states the sale of alum baking powders is prohibited by law. All baking powders contain starch as a filler. From what has been said it is clear that *baking powder is a chemical preparation which when brought in contact with water liberates carbon dioxide.* The result of the action is similar to that of yeast, only much more rapid.

Indian corn *forms a considerable part of the food of all of the people of the United States. In the south corn bread is the only bread ever tasted by thousands.* Corn is also used extensively as a source of **corn starch**, from which corn sirup is made by boiling the starch with dilute sulphuric acid. After separating the starch from the grain, the residue is sold as a high

protein stock feed under the name of **gluten feed.** It contains from 24 to 28 per cent protein. Barley is chiefly used as a human food in the form of either **pearl or patent barley.** The former consists of the whole grain which has been polished after removal of the husk. The patent barley is simply pearl barley ground to a flour. *Barley is extensively used in the brewing of malt liquors.* In **brewing,** the seed is first allowed to sprout; thus starch is changed to sugars through the action of **diastase,** a ferment which occurs in germinating seeds. In turn the sugars are broken by yeasts into carbon dioxide and alcohol. The alcohol is the constituent wanted in these liquors. The residues from the brewing and distilling industries are dried. They are rich protein cattle foods, such as **brewers' grains** and **distillers' grains,** having a content of 25 to 30 per cent of protein. "Ajax" is such dried distillers' grain. The cereal grain **rice** is extremely rich in carbohydrates, and the common product on the market has been prepared by polishing the original seed. This polishing process was formerly carried out by rotating the seeds in drums lined with sheepskin, but is now done by agitating the grains with talc.

In preparing **breakfast foods,** the entire clean grain is ground or pulverized in some cases, while in others the bran and germ are first removed. In order to improve their keeping qualities, these foods are often sterilized before being put up in sealed packages. "Shredded wheat" is a preparation of whole wheat, which has been cooked, then shredded, and finally baked. "Force" is malted whole wheat in the form of flakes, cooked with steam. By malting or partly predigesting the starch, a part of the latter is made soluble. "Grapenuts" is another malted preparation of the entire wheat berry. The process of malting has changed some of the starch to sugar which gives the food its sweet taste.

Oatmeal is the whole oat with the outer husk removed. This reduces the cellulose content very materially. **Corn flakes** consist of cooked corn, which has been treated with malt honey, dried, rolled, and baked. *When the amount of food which a package of breakfast food contains is compared with its cost, the material is found to be quite expensive.* **The cereals had better be bought in proper form and cooked at home.**

The pulses *include beans, peas, and lentils. The characteristic of these leguminous foods is their richness in protein.* When their nutritive value is compared with their price, they are undoubtedly *the cheapest of all foods*, and on this account they may well be called the " poor man's beef." In an air-dried condition, they contain about 10 per cent of water, 20 to 25 per cent of protein, 60 per cent of carbohydrates, 1 to 2 per cent of fat, and 3 to 5 per cent of ash. The pulses are well supplied with carbohydrates, but they are poor in fat. The **peanut**, although botanically one of the pulses, really resembles the true nuts in composition, and like them it *is rich in protein and fat.*

Just as the larger portion of the grain of cereals is really a storehouse of nutrients for the use of the young plant, so **roots and tubers** may be regarded as reserves for the nourishment for the adult plant itself. During the prosperous days of spring and summer the plant lays by a goodly supply of food material for the seed production of the late autumn or the next spring. The **potato** is one of the standard vegetables, chiefly *valued because of its richness in starch.* There are produced in the United States annually about 300,000,000 bushels of potatoes. A fresh potato contains 82 per cent of water, 16 per cent of starch, and but 0.3 per cent of protein. The rest is ash and fat. From this it is clear that the potato is one-sided in composition and not like the pulses or the

cereals. A boiled potato contains less water than a fresh one, the content being about 75 per cent of water and 20 per cent of starch. There is some loss of material in boiling, but this loss consists mainly of mineral matter and proteins, and it is consequently slight.

The **turnip** is a vegetable of but slight nutritive value. It contains about 90 per cent of water and 5 to 8 per cent of carbohydrates. *Its water content is therefore even greater than that in milk.* None of this carbohydrate material is starch; it is present in the form of *pectins*. *The peculiar flavor of turnips which is so highly esteemed by some persons is due to complex sulphur compounds.* **Carrots** are decidedly more nutritious than turnips, mainly because of their richness in sugar, of which they contain nearly 10 per cent. On boiling carrots, about 25 per cent of this sugar is lost. **Cabbage** is a favorite vegetable with many people, especially those of the German race. It is also prized by good shepherds as a sheep feed. Cabbage contains 90 per cent of water, 2 per cent of protein, and 6 per cent of carbohydrates.

A distinguishing feature of roots and tubers is that they are rich in ash, especially in salts of potassium and calcium. This gives them greater value in diets and animal rations than they would otherwise deserve. The **rutabaga**, **radish**, **onion**, and **kohl-rabi** have compositions that are very similar to that of the cabbage. **Beet** roots are richer in sugar, containing from 10 to 12 per cent of readily available carbohydrates. The **sugar beet** is one of the main sources of our table sugar; it contains from 10 to 20 per cent of sucrose, cane sugar.

When judged strictly from a chemical standpoint, **fruits** would seem to be of little value as food, owing to their low content of the ordinary food constituents; but when the delight which they afford to the palate and the eagerness

with which they are eaten by normal children are considered, fruits certainly appear to be a very important human food. Fresh fruit in general contains about 85 per cent to 95 per cent of water, 0.5 per cent of protein, 0.5 per cent of fat, and 5 to 10 per cent of carbohydrates; also a small amount of ash and some organic acids. *Calculated on the amount of dry matter, the quantity of ash in fruits is really considerable.* The value of fruits as food depends largely on their mineral content and their flavor. The ash of fruits is rich in potassium, calcium, and iron salts, all of which are valuable to the animal cell. One-half of the carbohydrate content of fruits usually consists of fructose, fruit sugar. In apples, apricots, and pineapples, on the other hand, cane sugar is present.

The process of **canning fruit** is only about one hundred years old. It was first used by a Frenchman named Appert, who put fruit into cans or bottles which he sealed. He then placed the full cans or bottles in water and boiled them for some time. In Appert's time, and indeed until recent years, it was generally thought that the oxygen of the air caused the spoiling or decomposition of the food. To exclude the air from the cans was the explanation of success in canning in his time. It is now known that *it is not the oxygen of the air which causes canned fruit and vegetables to spoil, but rather the bacteria and other microscopic organisms that are present.* Appert's theory was wrong, but his method of sealing and cooking by which he really sterilized the contents of the cans was correct. If food is sterilized perfectly, and the opening of the jar is then plugged with sterile cotton, the contents will not ferment; for the bacteria and yeasts to which such changes are due cannot pass through the cotton.

Bacteria and yeasts *are in the air, in the soil, and on all vegetable matter.* In fact they are of almost universal occurrence. Some of them do great harm; but it is believed that

most of them are beneficial. *Fruits spoil because of the development of these microörganisms,* which bring about chemical changes in the tissues, as the disagreeable odors and flavors that are produced clearly indicate. *Bacteria grow luxuriantly in foods that are rich in nitrogenous material,* especially if the foods are warm and moist. Among such foods are meat, fish, eggs, peas, beans, and milk. All of these are difficult to preserve because of the ever present bacteria. *Thus it is that in warm, muggy weather fresh meat, fish, milk, etc., spoil quickly.* Bacteria do not develop in material containing large amounts of sugar; but when small quantities of sugar are present, both yeasts and bacteria grow readily. Fruits are usually slightly acid, and so in general they do not support bacterial growth. It is for this reason that *canned fruits are more commonly fermented by yeasts than by bacteria.* Some vegetable foods contain so much acid and so little nitrogenous material that very few bacteria or yeasts will attack them. Lemons, cranberries, and rhubarb belong to this class.

The proper use of higher temperatures is by far **the best way of checking the growth of microörganisms,** for most of them are destroyed if exposed for ten or fifteen minutes to the temperature of boiling water, *i.e.* 100° C. or 212° F. But *if the bacteria produce spores, boiling must be continued for an hour or more to insure their complete destruction. Yeasts and their spores are more easily destroyed than bacterial spores.* Of course, one generally does not know what kinds of organisms are present in the particular food that is to be canned, but it is well known that most fruits are acidic and do not favor the growth of bacteria, and that as a rule the yeasts can be destroyed by heating from 10 to 15 minutes at a temperature of 212° F. *If no living organisms or spores are left, and the sterilization of all the dishes and appliances used has been*

thorough, there is no reason why the fruit, if properly sealed, should not keep for a year or longer. When **fruit is preserved with a large amount of sugar** (a pound of sugar to a pound of fruit), it does not need to be hermetically sealed to protect it from yeasts or bacteria, because the thick sirup formed is not favorable to their growth. **Molds**, however, will grow on fruit when thus preserved, and it is better to seal the jars in any case.

Every housewife is familiar with molds, which under favorable conditions of moisture and warmth grow upon almost any kind of organic material. For example, *in damp, warm weather molds form in a short time on all sorts of starchy foods, such as boiled potatoes, breads, mush, etc., as well as on canned and preserved fruits.* **Molds develop from spores which are always floating about in the air.** When a spore falls upon a substance containing moisture and suitable food, it sends out fine threads, which branch and work their way over and into the attacked substance. If one of these spores drops on a jar of preserves or a tumbler of jelly, it will germinate, if there is moisture enough present and the temperature is favorable. Molds do not penetrate deeply into solid foods, or into liquids or semiliquids, because they require free oxygen for growth; but if given time, they will at ordinary room temperature gradually penetrate solid substances which contain moisture. For these reasons they will not cause the ordinary fermentation of canned fruit. *To kill mold spores, food must be exposed to a temperature of from 150° F. to 212° F., after which it should be kept in a cool, dry place and covered carefully so that no floating spores can find lodgment on its surface.*

Dates, figs, prunes, and other **dried fruits** generally contain from 20 to 25 per cent of water and 65 per cent of carbohydrates; of the latter a large part is sugar. The remainder

of these fruits is made up of ash, fat, and protein. They are valuable food materials, which are rich in alkali salts, and commonly act as mild laxatives.

Nuts *are among the most concentrated of foods* and therefore present a strong contrast to the water-containing fruits. *Nuts are rich in oil, carbohydrates, and proteins,* and this makes them extremely nutritious. Bulk for bulk, dry nuts are doubtless among the most nutritious foods in existence. Their general composition is roughly as follows: water, 4 to 5 per cent; protein, 15 to 20 per cent; fat, 50 to 60 per cent; carbohydrates, 9 to 12 per cent; cellulose, 3 to 5 per cent; mineral matter, 1 per cent. As the figures show, fatty matter predominates in nuts; and *no other vegetable substances,* unless they be flaxseed, cottonseed, and castor beans, *are so rich in fat as nuts.* Occasionally substitutes for ordinary butter are made from the expressed fats of nuts, as for example " nut butter " and " nuttolene."

It is not the nutritive value of **coffee, tea, and cocoa** that makes these beverages so important, but rather the peculiar stimulating and restorative effects which they produce. All of them owe their peculiar virtues to the alkaloids which they contain (see Chapter IX). In coffee and tea the effect of the alkaloid is more striking; but in cocoa it is masked by the heavy nutritious fats and carbohydrates that are present. The three alkaloids, caffeine, theine, and theobromine, contained in coffee, tea, and cocoa respectively, have a certain relation to one another, and consequently their effects are somewhat similar. The amount of these alkaloids present is not large, being only from 1 to 2 per cent; but like certain other medicines, such *alkaloids in even very small quantities readily affect the nervous system.* The coffee bean, the tea leaf, and the cocoa seed all come from plants that are grown in the tropics. Children ought not to drink either coffee

FIG. 92. — A coffee plant in blossom.

or tea because of their stimulating action; but they may use cocoa, for it is rich in the ordinary nutrients.

Spices, *such as* **pepper, allspice, nutmeg,** *and* **cinnamon,** *are not foods in the strict sense of the word,* but nevertheless they are essential constituents of the diet. They arouse the appetite and promote the secretion of the gastric juice. Their active principles and flavors depend upon the presence of volatile oils. These oils are not of the same kind as the oils of the ordinary seeds and the fats of meat, but they are other complex compounds which are volatile, as their name implies. **Vinegar** *is a condiment* in which the active principle is acetic acid. Vinegar is made by allowing the expressed juice of apples or other fruits to ferment. In this process the sugars

FIG. 93. — A tea plantation.

in the fruits are first changed to alcohols by yeasts and so the juice becomes **hard cider**; the alcohols are then changed to acetic acid by bacteria. These changes are brought about at ordinary temperatures, but they proceed more rapidly at 90°–100° F.

The **foods of animal origin** include meat, gelatin, fish, eggs, and milk. *These are all high protein materials.* **Meat** is characterized by a high ash, fat, and protein content. The amount of fat in meat varies greatly with the condition of the slaughtered animal. *Lean sirloin* has about the following composition: water, 71 per cent; protein, 25 per cent; fat, 3 per cent; ash, 1 to 2 per cent. A *fat sirloin* may contain as much as 30 per cent of fat. It should again be recalled that *there are no carbohydrates in meat.* Only fresh **liver** and **oysters** contain the animal starch called *glycogen,* and even this is present only in small quantities. A piece of *boiled meat can easily be torn into long, stringy fibers.* On microscopic examination these are found to be made up of bundles of

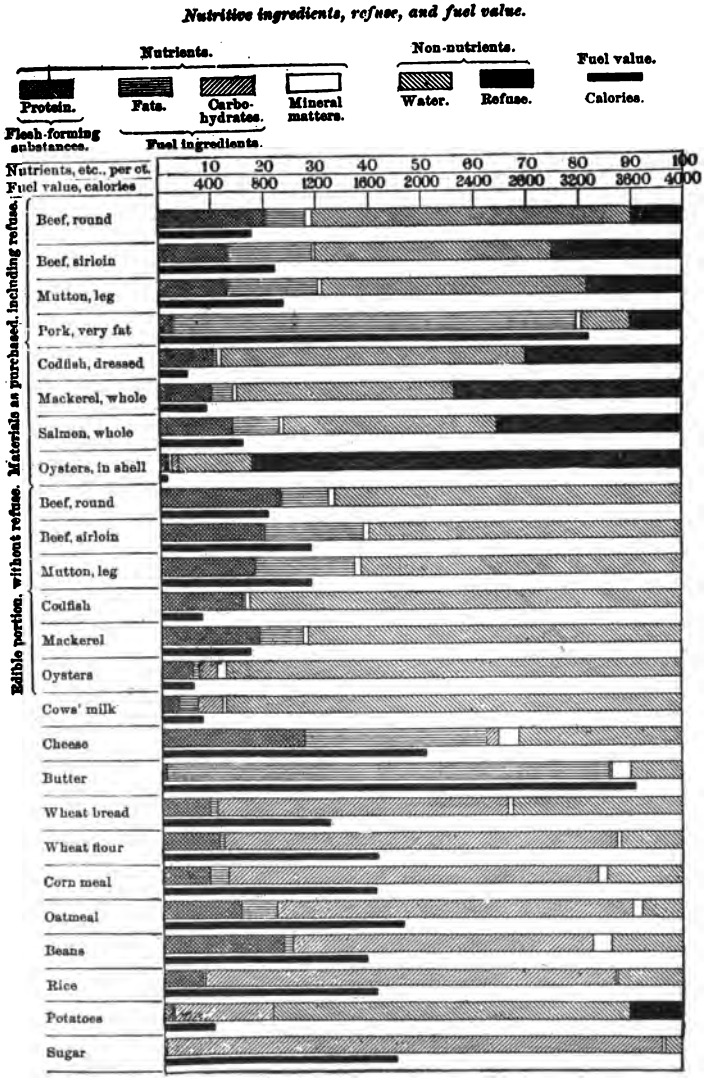


FIG. 94. — Composition of food materials.

microscopic tubes called **muscle fibers**. These fibers vary in length in different kinds of meat. In some meats they are short, as in the breast of chicken; in other cases they are much longer, as in the leg of the crab. The shorter these fibers are, the more tender and more easily digestible is the meat. *Meat should be cut across the grain*, that is, at right angles to the axes of the fibers, for then it is easier to chew it; and since the contents of the tubes are better exposed, the meat is also more completely digested. **Gelatin** *is the chemical basis of all jellies prepared from animal materials.* It is a protein and comes from the connective tissue and bones of animals. On boiling these tissues the gelatines are dissolved and on cooling the resulting solution sets to a jellylike mass. **Glue** *is crude gelatin.* It is commonly made from hide clippings. The purest form of gelatin is **isinglass**, which is made from the swimming bladders of fish, especially from those of sturgeons. Ordinary gelatin contains 85 per cent of protein and 12 to 13 per cent of water; but the common jelly prepared for the table at home contains only from 2 to 3 per cent of gelatin.

Fish, *like meat, contains protein, fat, and little or no carbohydrates.* In localities near the sea, fish is the cheapest source of protein. The composition of the different food fish is somewhat variable, but the water content is usually about 60 per cent, the remainder being bones, protein, fat, etc. The protein content varies from 10 to 16 per cent.

Next to milk, **eggs** *are the most wonderful foodstuff in the world.* Just as milk contains the various nutritive constituents required by the newly born young, so in the egg there are present all of the compounds that are necessary for the new life that lies dormant within its jellylike mass. Out of it may come bone, hair, blood, muscle, and feathers. In chemical language, *the egg contains an abundance of salts, proteins, and fats, but no carbohydrates.* The average weight

of a hen's egg is two ounces, about 11 per cent of which is shell, 57 per cent is white, and 32 per cent is yolk. The shell is carbonate of lime. The white is nearly a pure solution of the protein albumen, and contains 12 per cent of the latter, about 87 per cent of water, and 1 per cent of ash. The yolk is more complex than the white. It contains from 15 to 16 per cent of protein, from 32 to 33 per cent of fats, 50 per cent of water, and 1 per cent of mineral matter. *Some of the fats of the yolk are of a special kind, called lecithins; these are rich in phosphorus* (see Chapter IX).

In addition to the grains and the various products prepared from them, which serve as food for both man and animals, there are the bulky foods called **forage** which are so common on the farm. *These are the hays, straws, cornstalks, and corn silage.* As a rule *these fodders are rich in carbohydrates and poor in proteins,* but they differ greatly among themselves, depending upon the kind of plant from which they are made. Thus the **hays** made from such plants as *timothy, redtop, and June grass are comparatively poor in protein and rich in cellulose or crude fiber.* They contain about 6 per cent of protein and as much as 30 per cent of crude fiber. The rest of the plant consists chiefly of available carbohydrates. These available carbohydrates, however, are not really starches and sugars; they are more like cellulose in character. In contrast to the hays made from the above-named grasses, those from the *legumes, such as the clovers and alfalfa, contain from 12 to 15 per cent of protein* and nearly as much crude fiber and available carbohydrates as the hays of the first-mentioned group. **An acre of alfalfa will produce 1600 pounds of crude protein while an acre of timothy will yield but 300 to 400 pounds.** This is the reason why the legume hays are such valuable fodders. They serve as sources of protein to supplement the cereal grains in the ration. The

cereal straws, such as those of oats, rye, wheat, and barley, are generally considered poorer feeds than timothy, redtop, and June grass hays. The **straws are rich in fiber,** often containing as high as 40 per cent, *and have but little protein,* namely, from 3 to 5 per cent. **Corn stover** — which is left,

FIG. 95. — An alfalfa field. The caps will protect the hay.

after the ears have been harvested — *has a higher feeding value than generally considered.* It has been found that it will nearly furnish a maintenance ration for cattle. When field-cured, it contains about 19 per cent of fiber, 30 per cent of other carbohydrates, 40 per cent of water, and 3 to 4 per cent of protein.

It is becoming common practice wherever corn is grown to cut it before the kernel is hard and glazed and store it in a *silo.* There are probably 100,000 silos in America, and 95 per cent of these are filled with the corn plant. **Silage** *can be made from beet leaves, beet waste, pea vines from the canneries, and to some extent also from clover and alfalfa; but corn is the*

FIG. 96. — A modern concrete silo. (289)

main plant used at present for this purpose. When harvested and cut, the material is blown into the silo and firmly packed so as to exclude as much air as possible. The material in the silo soon becomes warm; the sugars of the plants break down to acids, alcohols, and carbon dioxide, and some of the proteins present are partly digested. The acids formed are mainly lactic and acetic, and the acidity may be as high as 1 to 2 per cent. *The formation of the acid keeps the silage from further decay.* There are some losses of dry matter in the making of silage. They have been estimated to amount to at least 10 per cent under the best conditions of handling. Nevertheless, these losses are probably smaller than if the same material were fed after it had been cured on the field. *Silage is relished by animals, and there is no waste when it is fed.*

QUESTIONS

1. What is the chief class of nutrients in vegetable foods? Name some foods in which this is an exception.

2. Name the cereal grains. From which of these is bread made?

3. What is the effect of baking on the starch granules? How does a very high temperature affect starches? How does it affect some of the proteins?

4. What makes bread dough light? What is baking powder?

5. What is gluten feed, and how does it differ from corn meal? What are "corn flakes" and "force"?

6. How do beans differ in composition from corn meal? Name a starchy tuber. How much water is there in fresh cabbage?

7. How much sugar is there in a sugar beet? What is the principal nutrient in a walnut?

8. What causes the spoiling of fruit? How can it be prevented?

9. What is the difference between a yeast and a mold?

10. Name some important foods of animal origin and state their chief characteristics. Of what is the shell of an egg composed?

11. How does timothy hay differ in composition from wheat flour? How does corn stover differ from alfalfa hay?

CHAPTER XX

MILK AND ITS PRODUCTS

DAIRY products have been used by the human race since earliest times, for *the oldest writers speak of milk, butter, and cheese.* At present *there is scarcely a person in the civilized world who does not consume milk or some of its products every day.* The dairy industry is consequently of very great importance. Its purpose is to furnish milk, butter, cheese, and various other products made from milk. Milk is usually produced on the farm, but occasionally large dairies are located in the city, where cows are stabled the year around and fed all of their feed in the manger. *Most of the milk consumed comes from cows, but in some countries goats' milk is much used.* Even in America goats' milk is prized for infant feeding by those in charge of infant hospitals.

Milk *is secreted from the blood.* In the case of the cow this function is performed by two large glands. While the blood is red, milk is never bloody unless the cow is sick. This shows that in making milk from the blood very great changes take place. The *milk is prepared from the blood by the cells of the glands.* The latter are rather soft, spongy organs that contain a fine network of ducts which are lined with secreting cells. These ducts, which lead to a small cistern located just over the teat, have their origin in clusters of active secreting cells. The gland is well supplied with blood, and *at no time does it contain much milk.* Most of the milk is actually formed while the cow is being milked, but just how this is really done is not understood.

Milk contains all the food that is necessary for the growth and development of young animals. It is a complete food, and through careful feeding, breeding, and selecting, man has developed the dairy cow so that she now yields much more milk than is needed for her offspring, for whom nature intended the milk. *Some cows have produced from ten to fifteen times their weight of milk in a year,* or over 20,000 pounds; but most of them give less than one-fourth of this quantity annually. *Milk is composed of the following substances: water, fat, casein, albumen, milk sugar, and ash.* A few other substances are present in small quantities, but they are of little practical importance. The average percentage composition of cow's milk is as follows:

Water	87.1 per cent		
Solids	12.9 per cent	Fat	3.9 per cent
		Casein	2.5
		Albumen	0.7
		Sugar	5.1
		Ash	0.7

It should be noted that the per cent of water is very high, being about the same as that in fresh vegetables.

The composition of milk *varies a great deal with different breeds of cows and with the individual animals.* The dairy breeds, such as *Jerseys and Guernseys, usually produce milk rich in fat, while the milk from Holstein and Ayrshire cows has a lower percentage of this constituent.* However, the breed of a cow is not a sure sign of the richness of her milk; more depends upon the cow herself. A cow gives her largest flow of milk after the calf is a few weeks old, but *the nutrients of the milk gradually increase in quantity as the period of lactation advances and as the flow decreases.* The first milk drawn from the udder is very low in fat, while the last fraction, called

stripper milk, will have as much as 10 per cent of fat. It is a common notion that the fat of the milk can be increased by the kind of feeds which a cow eats. But many well-con-

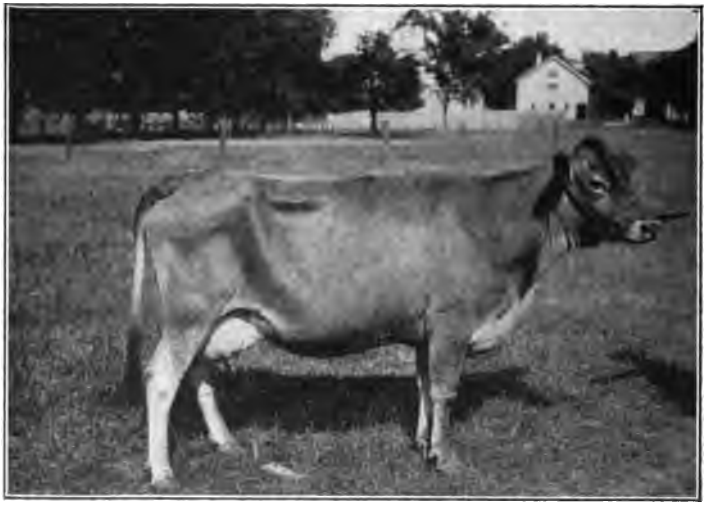
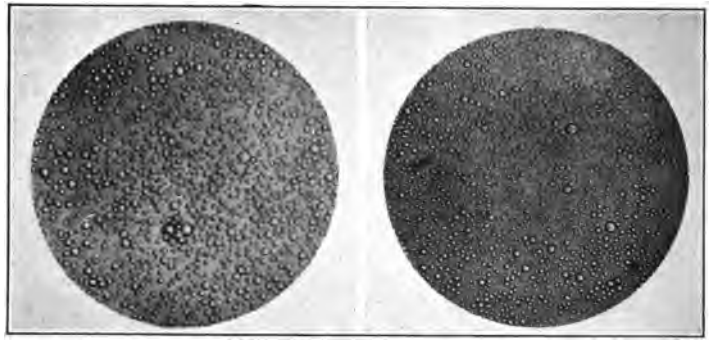


FIG. 97. — A pure-bred Jersey cow.

ducted experiments have proven that *such influence is very slight, if indeed the composition of milk can thus be affected at all.*

The fat *occurs in milk in the form of small globules.* They are very minute in size and it would take about 6000 of them placed side by side to make an inch. In the milk of Guernsey and Jersey cows the average size of the globules is considerably larger than in Holstein milk. This explains why the cream layer is formed so much more rapidly in Guernsey and Jersey milk than in Holstein milk. The globules of any sample vary greatly in size; and the largest are recovered in the cream when the milk is set, or run through a cream separator, while the smallest ones remain in the skimmed milk.

Casein *is the principal nitrogenous body in the milk* and gives to the skimmed milk its bluish white color. It can be separated from the milk by the addition of an acid, or by the action of rennet. In the **souring of milk**, during which process lactic acid is formed from the milk sugar by bacterial action, the casein is changed into a firm curd. To-day large amounts of skimmed milk are treated with sulphuric acid. The



(*A*) (*B*)

FIG. 98. — The size of fat globules compared. (*A*) Jersey milk. (*B*) Holstein milk.

precipitated casein is then partly dried and used in the manufacture of knife handles, billiard balls, and countless other articles. *When rennet is added to milk, the casein separates,* and the fat is firmly held in the curd. This forms the starting point in the manufacture of most forms of cheese. **Junket tablets** are dried rennet. The **albumen in milk** is, of course, also a nitrogenous compound. It is probably of use to the growing calf, but is not used commercially as a separate product. It may be obtained by boiling the **whey** which is left after removing the casein.

Milk sugar is a carbohydrate which has already been described (see Chapter IX). It is obtained from the milk

by evaporating the whey left in cheese making. *This sugar gives to milk its sweetish taste.* When separated from whey, it is used in the preparation of infant foods, and also as a vehicle or carrier for certain medicines. *It is the principal constituent in whey, and gives the latter value as a feed for pigs and calves.* The **ash of milk** is the mineral matter that is left after burning off the organic matter. It contains all of the substances that are necessary to build the bony structure of the growing animal, and furnishes the needed mineral matter for the formation of new tissue. The first milk secreted by the cow after the birth of the calf is yellowish and rather thick. This is called **colostrum** and is very rich in albumen, often containing as much as 13 per cent. *This first milk helps to start the action of the digestive tract of the young calf and is consequently very important to the latter.*

To produce milk which is pure and fit for drinking purposes, or for butter or cheese making, requires more care than any other work on the farm. Careful attention must be given to *the surroundings of the cow,* to *the actual process of milking,* to *the dishes in which the milk is placed,* and to *the place where it is kept.* When dust gets into milk, it carries with it small one-celled bodies, called bacteria. These are a form of plant life; they have been mentioned in previous chapters. At favorable temperatures these bacteria grow rapidly. In 24 hours a single bacterium may increase in numbers to several millions, which are all so small that they are not noticed except with a powerful microscope. However, *their presence and growth* **causes the milk to sour,** *for they change milk sugar to acid.* This acid is principally **lactic acid,** and it causes the milk to curdle. The common notion that thunderstorms cause milk to sour is entirely erroneous.

Disease-producing germs *or bacteria are often introduced through the dirt which gets into the milk. Scarlet fever, tuber-*

culosis, typhoid fever, cholera, and sore throat may be carried and spread in this way. The **house fly** is a great carrier of dirt and disease and should be kept out of the milk. Milk, too, has the power of absorbing gases and odors and readily acquires these from the air. A strong silage odor in the barn or onion odor in the refrigerator may be absorbed by the milk and make it objectionable as a drink. A foul-smelling stable will also taint the milk. *All of these facts emphasize the necessity of cleanliness in milk production* and the adoption of methods to check bacterial growth. The growth of bacteria can be checked by cooling the milk as soon as it is produced. This prevents a rapid development of these organisms, but it does not entirely check their growth. Nevertheless, *the sweetness of the milk is prolonged by this cooling process, which is a very practical method and the one usually followed.* **To destroy the organisms in milk** *either heat must be employed or chemical substances — antiseptics — must be added to the milk.* If the latter is heated enough to destroy all organisms (which requires a temperature above 100° C., *i.e.* 212° F.), it will turn brown and acquire a burned taste. The process of thus *destroying the organisms by heat is called sterilization.* To avoid its disadvantage, the process known as **pasteurization** is often substituted. In the process of pasteurization which was originated by the famous French chemist, Pasteur, and is named after him, the milk is heated for twenty minutes at only 60° to 80° C.; thus the flavor is but slightly affected, and yet most of the active bacteria are killed and the keeping

FIG. 99. — A dirty fly. Greatly magnified.

qualities of the milk are materially increased. By adding various substances to milk the growth of bacteria can be prevented. **When, however, antiseptics are added in sufficient quantities to prevent bacterial growth, the milk becomes unfit for human consumption.** The chief *preservatives in common* use are boric acid, salicylic acid, formaldehyde, and benzoic acid. Their use in any quantity is generally prohibited by law; and this is right, because *when used they show that uncleanly methods of milk production are being practiced.* **Milk produced under clean, sanitary conditions, and well cooled, will keep as long as it is necessary for transportation to the city and the consumer.**

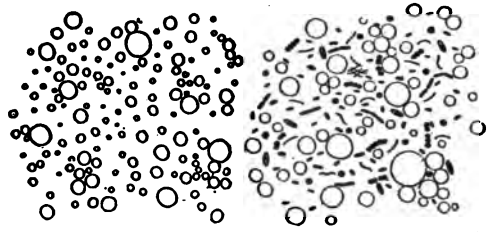


FIG. 100. — Pure and impure milk. The long black rods are bacteria.

The fat of milk exists in globules which tend to rise to the surface when the milk stands. After the fat has been removed, what is left is called **skimmed milk.** **Cream** is milk with a large amount of fat in it. *Cream can be separated from milk by gravitation, or by the much greater centrifugal force produced by rapid rotation in the centrifugal separator.* There are two ways of separating by gravity, namely, the *shallow-setting and the deep-setting methods.* In the **shallow-setting,** the milk is placed in pans two to four inches deep, cooled to 60° F., and allowed to stand from 24 to 36 hours. The cream is then removed with a slightly concaved ladle. Not all the fat is obtained in this way; indeed, only about 80 per cent of the fat is usually thus removed. In the **deep-setting** process, the milk is put into cans twenty inches deep

and less than one foot in diameter, which are then placed in ice-cold water. The cream rises rapidly and the operation is practically complete in 12 hours. In this way 90 to 95 per cent of the fat can be removed, depending upon the conditions of cooling, manipulation, and the breed of the cow whose milk is thus treated. *Because of its small globules, the fat of the Holstein milk cannot be as completely removed as that of the Jersey milk by this process.* **Cream separators** are machines which are now in general use. In them the milk runs into a bowl which is revolving several thousand times a minute. This revolving tends to throw the heavy particles to the outside of the bowl. Since the fat is not as heavy as the other constituents of the milk, it tends to come to the center of the bowl. The cover of the bowl is so constructed that the cream, or fat, can escape from the center of the cover, while the skim milk escapes from the opening near the edge of the bowl. *There are many different makes*

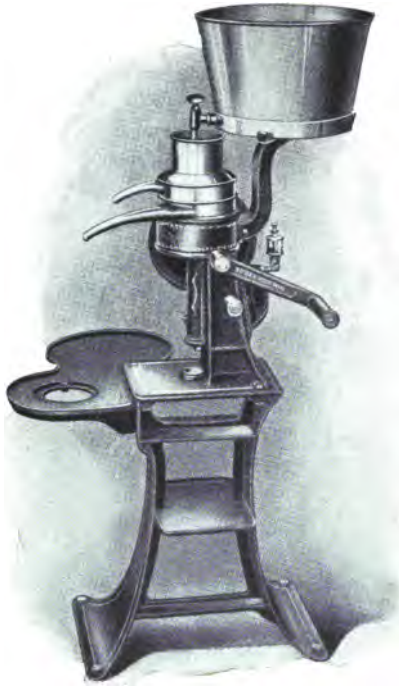


FIG. 101. — A modern cream separator.

of separators, but the principle of operation in all of them is practically the same as that just described. These machines recover from 97 to 98 per cent of the fat in the milk. A well-operated separator rarely leaves as much as 0.1 per cent of fat in the skimmed milk. *A good cream usually contains from 25 to 30 per cent of fat.*

Skimmed milk varies in composition according to the more or less complete removal of the fat. It differs from whole milk only in its smaller content of fat, and so *it still contains a valuable amount of foodstuffs* which should be used on the farm. *It is excellent for feeding pigs.* Though poor in fat, machine-separated milk has the advantage of being sweet and of keeping better than the product from other processes of skimming.

Upon agitating cream or milk for some time, the fat globules coalesce and **butter** separates out in irregular masses. **Churning** is a mechanical process in which the fat globules of the cream collide and adhere; the large, irregular masses thus formed then become centers of growth to which still other fat globules adhere. *To be of good quality, butter must possess a certain texture and grain and be neither hard nor smeary.* This desirable result can be secured only by churning at a favorable temperature. If the temperature is too low, the butter will be long in coming and very hard. If the temperature is too high, the butter will come quickly, but it will be greasy and destitute of grain. The temperature of churning should not vary more than from 45° to 65° F.; indeed, in most cases from 50° to 60° F. is chosen as the proper range. In **the manufacture of butter**, the cream is either allowed to sour (ripen) before it is churned or is churned directly without ripening. The former method yields the ordinary **market butter**, the latter method the **sweet cream butter.** In the ripening, acids and other products are formed, which give the butter a high flavor. Butter made

from sweet cream is usually not salted. It contains somewhat more water, fat, and casein than sour-cream butter. Ordinary market butter contains about 12 per cent of water, 84.2 per cent of fat, 1.3 per cent of curd, and 2.5 per cent of salt. Under the Federal Pure Food Law, butter must contain 82.5 per cent of fat and not more than 16 per cent of water. During the working of the butter a part of the **buttermilk** is squeezed out. This buttermilk does not vary much in composition from the skimmed milk.

Oleomargarine, which is much used as a substitute for butter, is made by churning together "oleo oil," "neutral oil," milk, and a small amount of cottonseed oil and peanut oil. "Oleo oil" is the liquid oil pressed out of beef fat. "Neutral oil" is melted lard. *Some of the oleomargarine is colored yellow so as to imitate butter, in which case it is taxed 10 cents per pound.* **Renovated butter** is made from old and rancid butter, by melting it, separating the fat from the casein, and blowing air through the fat to remove the unpleasant odors which are due to volatile acids that have formed. The liquid fat is then churned with milk, and a granular mass is obtained upon cooling. This is then worked, salted, and made up as butter.

Numerous kinds of **cheese** are made from milk. *All of the casein, nearly all of the fat, and also a large part of the ash of the milk are present in cheese.* The albumen and milk sugar are in the whey. Only two types of cheese which are commonly made in this country, namely, Cheddar cheese and cottage cheese, will be described here. **Cheddar cheese** is made by adding a small amount of rennet extract to the milk. The **rennet** is made by extracting the fourth stomach of the calf with a dilute salt solution. Rennet contains a ferment which causes the milk to curdle. In the process of Cheddar cheese making, the milk is warmed to 84° F. and "ripened"

to 0.25 per cent acidity. Rennet is then added, and when the curd is firm, it is cut into small cubes. In this process the fat has become entangled mechanically in the curd. The vat is now warmed, which causes the curd to shrink and harden. After being maintained warm for one to two hours, the whey is drawn off and the curd is piled. In this condition it mats into a solid mass. This is now passed through a grinding mill,

FIG. 102. — A cheese curing room.

salted, and pressed into molds. *The cheese is next put into curing rooms* at a temperature of from 50° to 60° F. and *allowed to ripen.* When three to four months old, it is put on the market. The practice of *ripening the cheese* at a temperature as low as 30° F. is coming into general use. The process of making **cottage cheese** is substantially as follows. Instead of adding rennet to curdle the milk, it may be allowed to stand until it is sour; or it may be soured by adding a "starter." A starter is a powder, containing living bacteria,

which will slowly produce lactic acid when added to milk. The curd thus obtained is strained from the whey, and after being salted, this makes sour-milk or cottage cheese.

The following data will give a general idea of the **composition of cheese.** Cured Cheddar cheese contains about 34 per cent of water, 35 per cent fat, 28 per cent casein and 3 per cent salt. **Swiss cheese,** which is made extensively in some parts of this country, has about the same composition. Some types of cheese, as **Camembert** and **Roquefort,** contain as high as 50 per cent of water. Milk is sometimes partly skimmed before being coagulated with rennet, and the product thus obtained is called **skimmed-milk cheese.** In many states its manufacture is prohibited.

There are still other dairy products on the market. **Condensed milk** is manufactured from whole milk or partly skimmed milk by evaporating off a certain portion of the water. The milk is evaporated in large vacuum pans to one-third or more of its original volume. Condensed milk should contain at least 8 per cent of fat. In many cases cane sugar has been added in large quantities. This aids in preserving the product, even after the cans are opened. A sweetened condensed milk may contain 40 per cent of cane sugar. Condensed milk is a good substitute for whole milk or cream, and it is used when fresh milk cannot be obtained, as on sea voyages, in mining camps, etc. **Milk powders** are made from skimmed or partly skimmed milk by allowing the milk to evaporate in a thin layer on a revolving drum. The product is scraped off in flakes, and placed on the market as thin, yellow scales. By a later process this powder is produced as a fine flour and this is destined to be a very efficient way of handling milk for long shipments. Most milk powders cannot be made from whole milk, as the high fat content of the powder impairs their keeping quality. **Ice cream** is made from rich milk or from

cream; the latter is the better. Ice cream consists of milk or cream, sugar, eggs, flavoring material, and sometimes cornstarch, gelatin, or gum tragacanth. The last three substances are added to give the ice cream smoothness and body.

No two things have done more to revolutionize the dairy industry than the cream separator and a simple test for fat.

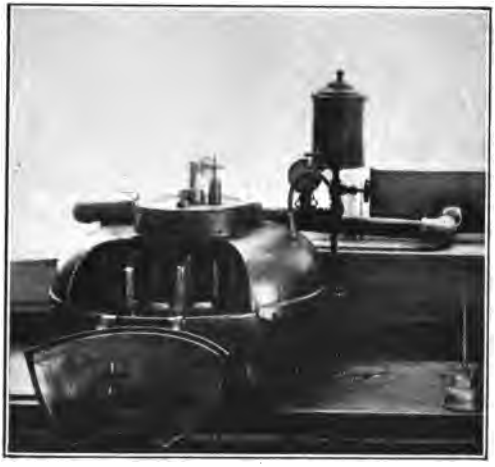


FIG. 103. — A Babcock tester run by steam.

The cream separator is a European invention, but the simple efficient fat test for milk was invented in America by Professor S. M. Babcock of the Wisconsin Experiment Station in 1890. **The Babcock test** *is a simple means of testing the milk for its amount of fat.* By its use the dairyman can determine which of his cows are paying their board, and in this way he can definitely ascertain the value of a cow for butter production. Knowing the per cent of butter fat in the milk and the quantity of milk produced, it becomes an easy matter to compute just how much butter fat is being produced. Again,

on the basis of the test, milk can be paid for on the basis of its fat content at creameries and cheese factories. Since the Babcock tester has become an established part of practically every dairy equipment, the old and but too common practice of watering milk has largely disappeared. Coupled with the Babcock test for the detection of water adulteration there should be used another equally important instrument, called the **lactometer**. This instrument resembles a floating thermometer (see Fig. 107). Its purpose is to determine the specific gravity of milk. Normal milk has a specific gravity of 1.029–1.034, and if the fat is partly removed, the specific gravity is raised; but if the adulterator now adds water, he may bring the specific gravity back to that of normal milk. The use of both tests detects such tampering with the milk. For a more just distribution of dividends at cheese factories, there is coming into use *a simple mechanical way of estimating the amount of casein in milk.* This is known as the **Hart casein test**; it was devised at the Wisconsin Experiment Station in 1907. This method is almost as simple as the Babcock test for fat, and it will help to solve the problem of proper payment for milk at cheese factories. *With both the fat and casein tests the cheese-yielding capacity of the milk can be determined.*

QUESTIONS

1. What does the milk of animals come from?
2. About how much fat is there in average milk, and how does it exist in the milk? What is this fat used for?
3. Which has the larger globules, Jersey or Holstein milk?
4. What is the principal use of casein in cows' milk?
5. How would you obtain the milk sugar from milk?
6. What precautions should be taken to produce clean milk? What is meant by pasteurization, and sterilization?

STEPHEN MOULTON BABCOCK. 1843–.
Scientist and inventor of the fat test for milk, which bears his name. This test has revolutionized dairying.

7. What is an antiseptic? Name two antiseptics and state whether they should ever be allowed in milk. Give reason for your answer.

8. Name three ways of separating cream from milk. How much fat is there in good cream?

9. Why does butter sometimes fail to come? How much water is there in ordinary butter?

10. Describe how oleomargarine and renovated butter are made.

11. How is Cheddar cheese made, and what is its general composition?

12. Who invented the commonly used fat test for milk, and where was it done? Of what importance is it? What two tests should be used at cheese factories for payment of milk?

13. What is a lactometer, and what is it used for?

CHAPTER XXI

POISONS FOR FARM AND ORCHARD PESTS

In order to have vigorous, healthy plants and sound fruits, the insect pests that are ever present to feast upon them and so impair their growth or even destroy them must be exterminated. To kill off such harmful insects that infest trees

Fig. 104. — Potato bugs and their eggs.

and other plants, various poisons have come into use. *These poisons are quite numerous, but they may all be divided into two great classes according to the type of insect that is to be destroyed.* Thus there are pests, like the potato bug, that feed directly on the whole leaf and will consequently be killed by any poison on the leaf. Such poisons which destroy because they are thus eaten by the insects are called *stomachic poisons.*

Again, there are other insects like the aphides or plant lice and the San José scale that get their nourishment not by eating the whole leaf, but by merely sucking out its juices. To exterminate pests of this kind requires a poison that acts

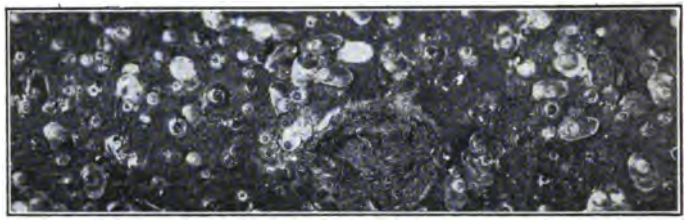


FIG. 105 (*A*). — San José scale on an apple twig, slightly enlarged.

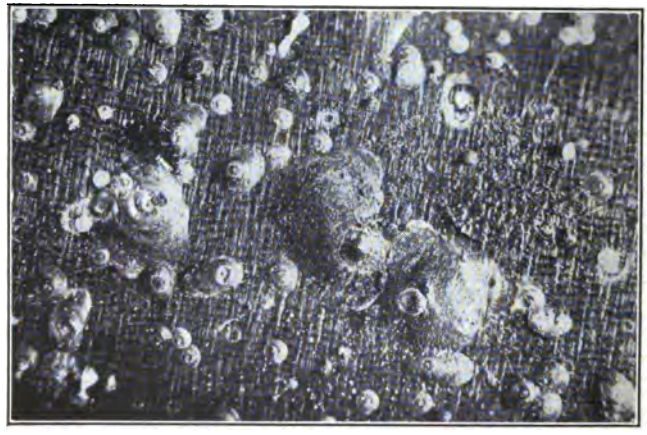


FIG. 105 (*B*). — San José scale on an apple twig, greatly enlarged.

by enveloping the insect and thus corroding it, poisoning it by absorption through its skin, or cutting off its opportunity to breathe. Poisons that act in this manner are called *contact poisons*.

Stomachic poisons generally contain arsenic, which is present not as the free element, but either as the oxide As_2O_3,

also known as white arsenic or arsenious acid, or as some more complex arsenical compound, like Paris Green or lead arsenate. The exact composition of these compounds has already been given in Chapter VII. **White arsenic** was first used as an insecticide; but though not copiously soluble in water, it nevertheless dissolves sufficiently to yield a solution that *corrodes or "burns" the foliage.* It was therefore soon discarded. In fact, *any arsenical poison that is to be used as an insecticide must have very little uncombined white arsenic,* As_2O_3, *in it.*

Paris green *is at present the standard poison for all insects that bite and swallow their food.* Its composition is expressed by the formula $Cu_3As_2O_6 \cdot Cu(C_2H_3O_2)_2$. This poison is prepared by adding a hot solution of arsenious acid to a hot solution of copper acetate. From this mixture Paris green separates out as a fine powder of a bright green color. Pure Paris green is almost insoluble in water, but *it will readily dissolve in ammonia water,* the resulting solution being dark blue. *This is a test by means of which certain impurities in Paris green may be detected.* Thus, if a sample of the substance is adulterated with gypsum, as sometimes occurs, the latter will form a white suspension in the ammonia water and finally settle out on the bottom of the glass. If no solids separate out, it of course only shows that impurities which are insoluble in ammonia are not present. *The glass test may also be applied.* In this test a small amount of the Paris green is placed in a glass; this is then inclined and gently tapped so that the poison will slowly move down the inclined plane. In the case of pure "green," the dust will be of a bright green color. If it is impure, it may yield a white, pale-green streak, depending upon the color of the adulterating substance that is present. Examination under the microscope will also help to determine its purity. Paris green contains about 58 per

cent of arsenious acid, 31 per cent of copper oxide and 10 per cent of acetic acid. These are all chemically combined in one compound, whose composition is indicated by the formula given above. *Paris green should not contain over 5 per cent of free arsenious acid, or it will seriously burn the foliage.* **For use in spraying,** *from 6 to 8 ounces of Paris green should be thoroughly worked into a paste with a little water, and then added to 50 gallons of water. In addition, about 2 pounds of lime should be added;* this will neutralize any free arsenious acid present, and also act as a "marker" on the sprayed plants.

London purple is an arsenical poison which was first imported from England as a substitute for Paris green. It is prepared by boiling with slaked lime a purple, arsenious acid bearing residue from the dye industry. Arsenite of lime is thus formed, together with some arsenate, which contains more oxygen than the arsenite. *London purple is more injurious to foliage than good Paris green,* because of its content of free arsenious acid. *It should always be used with lime.*

Lead arsenate was recommended as an insecticide in 1892 and was first used against the tent caterpillar. It is prepared by adding lead acetate solution to sodium arsenate also dissolved in water. These substances dissolve readily in the cold and react to form sodium acetate and lead arsenate $Pb_3(AsO_4)_2$. *This poison should be handled as a paste, for when once dried it does not again remain well in suspension.* It is the most insoluble of the insecticides now in use. Furthermore, *it adheres very tenaciously to the leaves and is least liable to scorch them.* **For spraying purposes** *two pounds of the commerical paste in 50 gallons of water is the proportion commonly used.*

Hellebore is often recommended as an insect poison. This is the ground root of the poke-root plant. **Pyrethrum,** or

insect powder, which comes from the flower heads of certain plants, is another poison of similar character. Both of these materials contain poisonous substances, but *they deteriorate much with age.*

Contact poisons may act (1) *by their caustic properties*, (2) *by poisoning because they are absorbed by the surface of the insect*, or (3) *by closing up the insect's breathing tubes.* Lime-sulphur and kerosene emulsion are types of these poisons. They are commonly used against the scale insects. **Lime-sulphur** was first employed in this country as a sheep dip, but in 1886 its use against the San José scale began. It is now either purchased in the commercial form or made at home. It may be prepared by heating together 80 pounds of high grade flowers of sulphur, 40 pounds of burned lime, and 50 gallons of water. A rather pure lime should be used, one that is low in magnesium, or otherwise there will be a waste of sulphur. The lime is first slaked in an iron kettle; and the sulphur is separately thoroughly mixed with a small amount of water and then added to the lime, after which the mass is boiled for 45 minutes. The mixture is best when boiled by passing steam through it. During the heating process, the lime combines with the sulphur forming calcium sulphides of variable composition. On standing, there results a yellow to orange colored solution, under which is an insoluble sludge. This

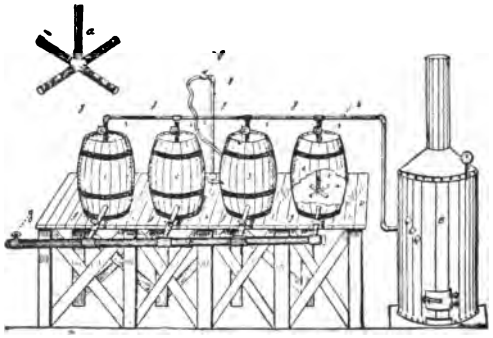


FIG. 106. — A simple home-made arrangement for making lime-sulphur.

consists of undissolved sulphur, calcium sulphite, and calcium sulphate. During the boiling, the solution should constantly be kept up to the 50-gallon mark by adding water as it evaporates. When finished, the clear liquid should be strained off and tightly barreled, because it does not keep in contact with the air, which oxidizes it.

For the purpose of exterminating scale insects the liquid is diluted to test 4.5° to 5.0° Baumé. For summer work, the solution is diluted with water to test but 1° Baumé. By this is meant that the strength of the solution must be so adjusted that a Baumé hydrometer (Fig. 107) will sink to the marks mentioned before the liquid is used. It is believed that on the tree the calcium sulphides decompose and leave free sulphur as the active agent.

FIG. 107. — A Baumé hydrometer in use.

Kerosene in the form of **kerosene emulsion** is also often used as an insecticide. Kerosene is a compound of hydrogen and carbon and is prepared from crude petroleum (see Chapter IX). It kills insects, and when applied to pools of stagnant water, it suffocates the emerging pupæ of mosquitoes. *When applied directly to plants, it kills them.* Kerosene cannot be diluted with water because it will not mix with the latter. For this reason kerosene must be mixed with some material that will carry the oil, as it were, and keep it from separating out from the mixture. For this purpose a soap is usually employed. *Kerosene emulsion may be made by dissolving one pound of soap in 2.5 gallons of water, then adding 2.5 gallons of kerosene to the solution, and thoroughly mixing by pumping the entire mixture through a bucket sprayer. Diluted to from 20 to 30*

gallons, the mixture may be used against scale and other sucking insects.

Tobacco decoctions can also be used as germicides; as such, their value depends upon the poisonous properties of the *nicotine* they contain. This alkaloid is soluble in water, and

FIG. 108. — Plant lice (aphis) on maple leaf.

so hot water extracts of the stalk and waste tobacco are made, cooled, and then used as an insecticide. *It is common practice to burn tobacco in greenhouses to get rid of certain plant lice.*

Gaseous insecticides are often used against insects that are particularly difficult to attack. Of these substances **hydrocyanic acid gas** is the most effective. **Cyanides are deadly**

poisons and should never be handled with the fingers (see Chapter IX). Hydrocyanic acid gas is made by treating potassium cyanide with sulphuric acid. Two ounces of concentrated sulphuric acid are carefully mixed with 4 ounces of water; this is placed in an earthenware vessel and then 1 ounce of potassium cyanide is added. This is the proper quantity for 100 cubic feet of space. **The gas is a deadly poison and one breath of it may be fatal; it should therefore by no means be inhaled.** To retain the gas about the plant and make its action effective, it should be applied in tightly closed rooms or buildings, or under tents, as is done in the treatment of San José scale on the orange trees of California. *The inclosure should afterwards be opened from the outside and thoroughly aired before it is entered by any one.*

Carbon bisulphide is a colorless, volatile liquid made from carbon and sulphur (see Chapter VII). The liquid is volatile and its vapors are fatal to insects. Being heavier than air, the vapors will settle through a mass of grain, and so *they are very effective in killing grain weevils.* A teaspoonful for each cubic foot of space, placed in a shallow dish upon the surface of the grain, will kill the insects present. The vapors settle down through the grain, and may later, after their work is done, be released by boring holes through the walls of the bin. *Ants, moles, prairie dogs, and similar pests are exterminated by placing cotton saturated with carbon bisulphide in the heaps or runs, and covering tightly with earth.* Carbon bisulphide should never be brought near a flame, for it is even more dangerous than gasoline.

Fungicides are materials for the destruction of certain parasitic plants called fungi. *The chief fungicide in common use is* **Bordeaux mixture**, which originated in France for the control of the downy mildew on grapes. It is made by suspending 4 pounds of copper sulphate (also called blue vitriol

or bluestone, see Chapter XI) in a gunny sack in about 15 gallons of water. The copper sulphate dissolves to a clear blue solution. Six pounds of lime are slaked and diluted to about

FIG. 109. — Students spraying an orchard.

15 gallons. The two solutions are then mixed and diluted to 50 gallons. This is the Bordeaux mixture. It is a suspension of rather insoluble salts of copper and lime of complex

composition. When applied to the foliage, the copper is slowly brought into solution by the action of the carbon dioxide of the air, and thus becomes active in destroying the pest. *The lime should always be in excess in Bordeaux, and sufficient in quantity so that the mixture will turn red litmus paper blue.* As long as copper sulphate is in excess, the reaction will be acid to litmus and the scorching of the foliage will follow. The potassium ferrocyanide test may also be used to detect an excess of copper sulphate in solution; for when Bordeaux mixture is filtered and a solution of potassium ferrocyanide is added to the filtrate, a brown precipitate of copper ferrocyanide forms if free copper sulphate is present. *Bordeaux mixture is a fungicide,* and it alone has no effective poisonous action upon plant insects. *To make it an insecticide, either Paris green or lead arsenate should be added.* This is now frequently done in the proportion of 4 ounces to 50 gallons of the Bordeaux mixture.

There are still **other poisons** that are sometimes used in the home and barn which must be briefly discussed. **Formalin,** *or formaldehyde, is a powerful disinfectant;* that is, it causes the complete destruction of disease germs. It is a product of the oxidation of wood alcohol, and is put on the market as a 38 to 40 per cent solution in water. For killing smut spores on grain, the seeds are immersed for ten minutes in a solution of one pint of the above 40 per cent solution to 20 gallons of water. It must be understood that *formaldehyde is never used as a spray for foliage.* For other purposes of disinfecting, as, for example, for use in water closets or for disinfecting sick rooms, it has no superior.

Mercuric chloride, corrosive sublimate, *is another common disinfectant* (see Chapter XI). It is put up in tablet forms with ammonium chloride to hasten its solution in water. One part in 1000 makes a sufficiently strong solution to kill

most germs. *Corrosive sublimate is a stomachic poison,* and forms insoluble compounds with proteins. Consequently an *antidote for it is raw eggs or milk.*

Chloride of lime or **bleaching powder** (see Chapter VI) *is a very common disinfecting and bleaching agent.* It is a powder prepared by passing chlorine into slaked lime. It always has a strong odor of chlorine. It is unstable and slowly gives up its oxygen, leaving calcium chloride behind. The oxygen thus liberated is the active agent which destroys germs and other organic matter.

Javelle water, which is also used for disinfecting purposes, is made by treating bleaching powder with washing soda and allowing the precipitate to settle. The clear solution is employed.

Lime *alone is one of the cheapest and most useful of the various disinfectants.* It will destroy organic matter as well as bacteria, and therefore it is useful in disposing of the bodies of animals which have died of some disease; it also serves as a whitewash in barns and pens. Sixty parts of water to 100 parts of lime will make a good whitewash or "milk of lime"; it *must be fresh to be active, for* **air-slaked lime** *is the carbonate and has no germicidal power.* When applied with a spray pump, whitewash will enter the cracks in stables, etc., and be very effective. *Whitewash, containing a little carbolic acid, is an especially good remedy in the poultry house for lice and vermin.*

The **burning of sulphur** *is one of the oldest methods of disinfecting;* but it is not very reliable. Sulphur dioxide destroys insects, vermin, and animal life in general, but it does not destroy the spores of bacteria. *In order to act at all, sulphur dioxide requires the presence of moisture.* Sulphur may be procured everywhere, and it has the further advantage that it is cheap. In **fumigating with sulphur** proceed as

follows: In a washtub put water to a depth of about four inches. In this, place several bricks or flat stones so that they will protrude above the water about two inches. Upon the bricks or stones, which should be in about the middle of the tub, place an old iron kettle. See that it rests securely on the bricks. Now place pulverized sulphur in the kettle, from 1 to 3 pounds, according to the size of the room to be fumigated. Open all the doors of closets, cupboards, etc., in the room. Also spread out the bedding. Now, having the washtub near the center of the room, insert into the heap of sulphur a piece of blotting paper about 3 by 8 inches, which has been soaked in alcohol. Wood alcohol or denatured alcohol or even strong rum or whisky will do. Mount this paper so that about 3 to 4 inches will be immersed in the powdered sulphur. Now with a match light the upper end of the paper. The alcohol will take fire readily and burn with a hot blue flame which will be communicated to the sulphur. As soon as the alcohol is burning well, leave the room and close the door tightly. After several hours, all the doors and windows may be opened so as to air the room well. It should not be inhabited till all odor of sulphur dioxide is gone. *In cities the health officer will attend to the fumigation of rooms that have been occupied by patients having contagious diseases.*

Certain products that are prepared from coal tar are also good disinfectants. Among these may be mentioned **carbolic acid, cresol, creosote,** *and* **naphthaline.** The latter is used in **moth balls. Lysol, carboleum, germol,** *and* **zenoleum** *are used in dips.* These are all coal tar products, as are also the so-called **fly removers** which are sprayed upon animals to protect them from flies.

Some of the advertised **lice exterminators** contain tobacco and sulphur as the active materials. These are not as ef-

fective as lime and carbolic acid. *In view of the fact that so many of these secretly compounded materials are fraudulent, ineffective, or very expensive, it is well to be sure of what the mixtures contain so as to be able to judge of their worth before purchasing them.*

It should be remembered that nature has provided good, efficient disinfecting agents, which, however, must be allowed to act freely. **Sunlight** *is a powerful germ killer.* If sunlight could be carried into every nook or corner of the house or the barn, there would seldom be any need of chemical disinfectants. *The germs of tuberculosis, diphtheria, and typhoid fever are killed by direct sunlight in six to eight hours.* The diffused sunlight which reaches our rooms is not nearly so powerful as direct sunlight, and it usually requires several days for it to destroy germ life. *Another common agent for killing bacteria is* **heat.** It may be employed dry, as by baking in an oven, or moist as by boiling with water or steam. Dry heat is not as effective in destroying germs as moist heat. All germs in clothing may be destroyed by boiling it for from five to ten minutes. *Floors and walls may be disinfected cheaply and efficiently with boiling water to which some lye has been added.*

QUESTIONS

1. Name two ways in which insects feed. How do the two classes of poisons used to exterminate insects act?

2. What is the common poisonous ingredient of the stomachic poisons? What is Paris green made of? Give two ways for testing Paris green for purity.

3. Why should lead arsenate be purchased as a paste?

4. How may lime-sulphur be prepared? Why should a high-grade lime be used?

5. For what purpose is kerosene emulsion used? Why is soap used in its preparation?

6. What is the substance used in killing ants, gophers, etc.? Why should flames be avoided when using this substance?

7. What is Bordeaux mixture, and why should lime be used in excess in its preparation?

8. What is formalin and when can it be used advantageously?

9. Name the chemical elements in bleaching powder. State how it acts.

10. Why is whitewash a good material to be used in stables and hen-houses?

11. What is the action of sunlight on germs?

CHAPTER XXII

PRACTICAL EXPERIMENTS TO BE PERFORMED IN THE LABORATORY[1]

LABORATORY MANIPULATION

Cutting glass tubing. Place the tubing on the desk, and firmly draw the edge of a triangular file across it two or three times (Fig. 110) at the place where the glass is to be cut. Then take the tube in the hands and with the thumbs placed near together, just back of the scratch, gently pull the glass apart, at the same time pressing outward with the thumbs (Fig. 111). If the tube does not break, make a deeper scratch. If the tube is large, it may be necessary to file around the glass. The broken edges will be sharp and should be rounded by rotating them in the tip of a Bunsen burner.

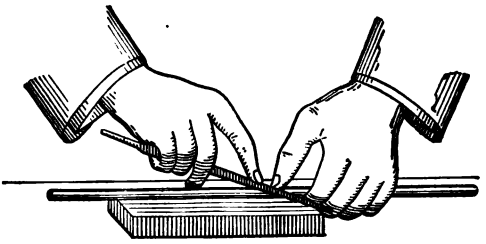


FIG. 110. — Cutting glass tubing.

Bending glass tubing. Use an ordinary gas flame, or the luminous Bunsen burner spread out by a "wing top." Hold the tube in the flame, as shown in Fig. 112. Rotate the

[1] Lists of apparatus and chemicals required for these experiments are given at the end of this chapter. The teacher will see to it that students use proper care in handling flames and dangerous substances.

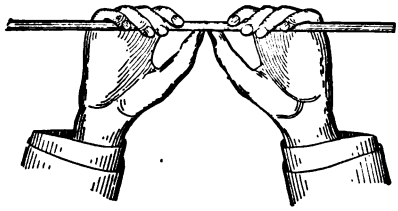


FIG. 111. — Breaking glass tubing.

tube and heat it until it begins to bend by its own weight; then remove it from the flame and carefully bend the tube to the desired shape. If the tube is heated uniformly and not too highly, the bore of the tube will not contract or flatten at the bend. (See Fig. 113.)

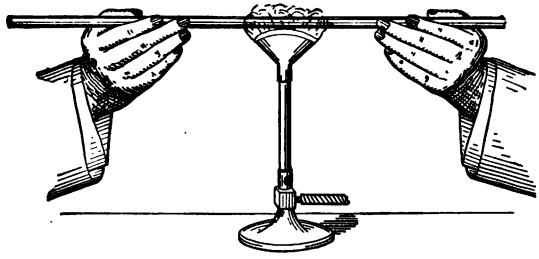


FIG. 112. — Bending glass tubing.

Fitting corks to glass tubing. Select a cork of suitable size for the test tube or flask to be used. Soften the cork by means of a cork press, or by rolling it between the desk and a block of wood. Select a cork borer slightly smaller than the tube to be inserted. Place the cork on the desk and bore through it,

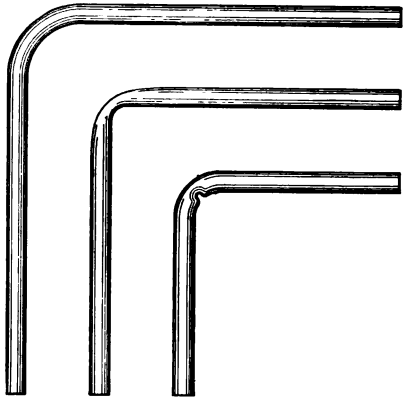


FIG. 113. — Good and bad bends of glass tubing.

not by merely pressing, but by rotating the borer under slight pressure, as in using a gimlet. Keep the borer upright and at right angles to the cork. If the hole is too small, a round file may be used to enlarge it.

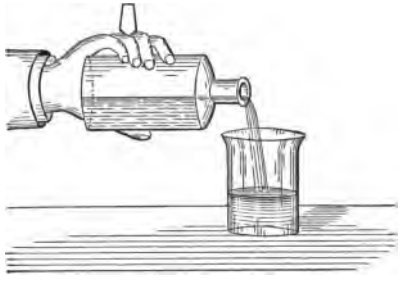


Fig. 114.— Pouring from a bottle.

Pouring liquids from bottles. Remove the bottle from the shelf, lift the stopper with the fingers, as shown in Fig. 114, and pour into the test tube. When the required amount of acid has been poured out, touch the neck of the bottle with the test tube to catch the drops on the edge and thus prevent them from streaking down the side of the bottle and

Fig. 115. — Pouring from a beaker.

on to the shelf or table. When pouring from one beaker into another, a glass rod, held as shown in Fig. 115, will prevent the liquid from running down the side of the beaker.

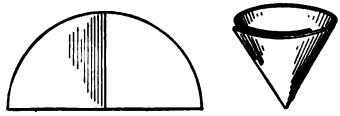


FIG. 116. — Folding a filter.

Filtering. This ordinarily has for its purpose the separation of a liquid from a solid. Place the funnel in the arm of a wooden stand. Fold the filter paper into halves and then at right angles into quarters (Fig. 116). The folded filter paper is opened so as to form a cone half of which is three thicknesses of paper and the remainder one thickness. Now fit the cone into a funnel of such a size that the edge of the paper does not quite reach the top. If the paper does not

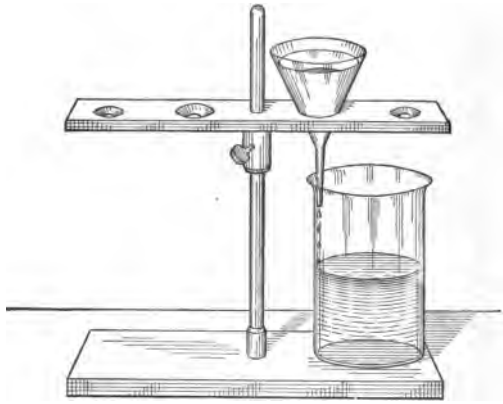


FIG. 117. — A filter ready for use.

fit, vary the folding slightly until it does. Now place the paper into the funnel and moisten it so that it adheres to the sides of the funnel. Press the paper gently with the fingers against the side of the funnel to remove any air bubbles. The filter is now ready for use (Fig. 117).

Heating liquids in test tubes. Fill the test tube only about one-third full. Hold it between the fingers and thumb, and constantly rotate it in the flame so as to apply heat uniformly. The heat should be applied to the upper end of the liquid and not to the bottom or to the glass above the liquid, as

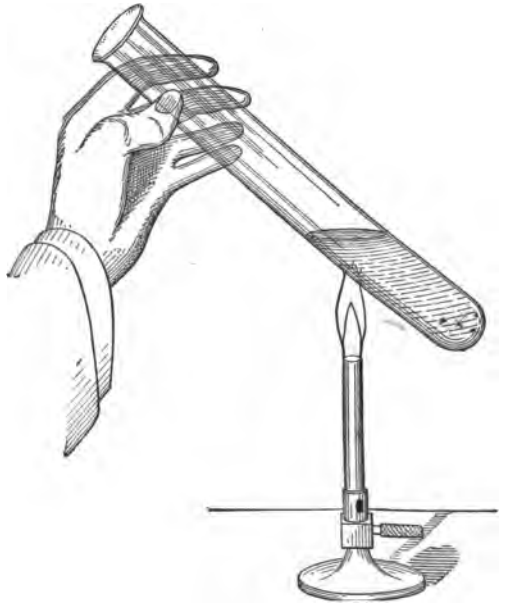


FIG. 118. — Heating a liquid in a test tube.

shown in Fig. 118. A test-tube holder or a band of paper may be used to hold the tube.

In the following experiments the student should record his observations and conclusions neatly and carefully in a laboratory notebook, which is to be submitted to the teacher for corrections and suggestions **upon completion of each experiment.**

EXPERIMENTS TO ACCOMPANY CHAPTER I

1. **Chemical and physical changes.** Examine a piece of rock candy. Notice its properties, such as its color, taste, hardness, and crystalline form. Describe its taste. (1) Grind a piece in a mortar. Has the taste changed? (2) Partly fill a beaker with water, and add a little of the powdered candy and stir with a glass rod. Taste the solution. Does the sugar still exist? (3) Heat 5 grams of the powdered candy in a dry test tube, applying the heat gently at first, and noting carefully the changes. When no further change takes place, even at a much higher temperature, allow the tube to cool. Taste the residue. Will it dissolve in water? Notice its color. What does the new substance resemble? Has there been a chemical change? (4) Heat a platinum wire or a piece of porcelain, and then let it cool. Was there a change in this case? Was it physical or chemical?

2. **Compounds and mixtures.** (1) Note the visible properties of a piece of sulphur. Test it with a magnet. Is it attracted? Drop a piece of roll sulphur, as large as a pea, in a test tube and add 5 to 10 cc. of carbon bisulphide and shake. **Use carbon bisulphide sparingly. Remember to have no flame in the room while working with it, for it is more dangerous than gasoline.** Does the sulphur dissolve? Pour off the liquid from the undissolved sulphur into a watch glass, and let it evaporate. Is there a residue on the watch glass? Is it like sulphur? (2) Examine some powdered iron. Test it with a magnet. See if it will dissolve in carbon bisulphide. (3) Stir together 3 grams of powdered sulphur and 5 grams of powdered iron. What is the color of the new powder? Bring a magnet to some of the mixture. Does the iron still exist? Put some of the powder in a test tube, add carbon bisulphide, shake, pour off the

liquid, and evaporate some of it on a watch glass. Does the sulphur still exist in the mixture?

3. **Chemical change.** Heat the rest of the mixture of sulphur and iron from the above experiment in a test tube with a small flame. When the mass begins to glow like a red-hot coal, remove the tube from the flame. Lay the tube down to cool. There is probably a little melted sulphur part way up the inside of the tube. Do not allow this to run down and mix with the substance at the bottom of the tube. When the tube is cool, break it open and examine the substance in the bottom where the glow had been. See if any sulphur can be dissolved out of it with carbon bisulphide. Test the black material with a magnet. Can you detect either sulphur or iron in the new substance?

EXPERIMENTS TO ACCOMPANY CHAPTER II

4. **Electrolysis of water.** This experiment is to be done by the instructor. Set up the apparatus as shown in Fig. 1. The water used should be slightly acidulated with sulphuric acid, for pure distilled water will not conduct the electric current sufficiently well. Three or four dry cells will suffice. Note the difference in the volume of gas formed in the two arms.

5. **Explosion of hydrogen and oxygen mixture.** This experiment is to be done by the instructor, who will prepare oxyhydrogen gas by electrolysis of water, and then pass the mixture into the eudiometer tube over mercury. The explosion is effected by means of the electric spark, the apparatus for this purpose being arranged as shown in Fig. 2.

6. **Salts in well water.** Place some ordinary service or well water on a watch glass, and set it in a warm place to evaporate. The evaporation will proceed best by placing

the watch glass on a hot water bath. Note the residue left in the dish.

7. **Preparation of distilled water.** Set up an apparatus as shown in Fig. 3. Place from 200 to 300 cc. of water in the flask. Apply heat gently, and collect 100 cc. of the distilled water. Be sure that cold water is running through the condenser. Observe the difference in the taste and odor of the water before and after distillation. Evaporate a few cubic centimeters of the distillate by gentle heat on a clean watch glass. Is there any residue left? How does this result compare with that of experiment 6?

8. **Water in various substances.** (1) Take some green leaves of any plant, also a carrot, or a potato to represent a root, and some wheat or corn as samples of seed. Weigh out 50 grams of each in weighed dishes of tinned iron or porcelain. Then put the dishes in a warm place and cover each of them with a glass plate. Window glass will do. In a few minutes the plates will be dimmed over with water that has been driven out of the plant by heat. *Save the materials for experiment 9.* (2) Heat pieces of beef, old mortar, and epsom salts each separately in a large test tube and observe the water that condenses on the side of the tube (see Fig. 4). The pieces to be heated should be of about the size of a small hickory nut.

9. **Amount of water in a plant.** When it has been seen that the plant tissue in experiment 8 is losing water, remove the glass covers and put the dishes in the open air or on the water bath to dry. This will require a day or so. On reweighing the dishes, it will be found that the contents have lost a good deal of water; 80 to 90 per cent in the case of the green leaves. The roots lose nearly as much, while the seeds lose only from 10 to 15 per cent. *Save the material for experiment 10.*

10. **Ash in a plant.** Take the residue from experiment 9 and heat it gently over a flame in a porcelain dish until everything possible has been burned off. There will remain a gray or white ash which may amount to from 2 to 5 per cent of the dry matter that was left after removing the water. *In the plant ash only a few elements are found, but these are the same whatever the plant, or wherever it was grown.* The elements in the ash are present as phosphates, sulphates, carbonates, chlorides, and silicates. The teacher will show the presence of calcium and phosphorus in the ash by simple qualitative tests.

11. **Deposits of lime by boiling water.** Boil for a few minutes about 200 cc. of hard well water or "White Rock" mineral water in a flask. After the water is cool, note any sediment of lime or turbidity of the water. Boiling has expelled the carbon dioxide, which has held the carbonates of lime and magnesia in solution. When this carbon dioxide is driven off, the carbonates precipitate. The crust in teakettles and the scale in boilers are chiefly of this nature. These deposits interfere with good heating and steaming.

12. **Hard and soft water.** Place about 50 cc. of very hard water in a 200-cc. cylinder or flask. This water may be prepared, if necessary, by adding 0.1 or 0.2 gram of calcium chloride to 500 cc. of ordinary water. Add to this a measured quantity of soap solution, made as stated below. Mix well, and notice how much soap solution must be added before a permanent lather is formed. Also notice the precipitate of lime soap. Repeat the experiment, using rain water and tap water respectively. Which of these waters requires the most soap solution? Hardness of water is due to the lime and magnesia salts in solution. *To prepare the soap solution,* scrape 10 grams of castile soap to fine shavings, and dissolve these in a liter of denatured alcohol; then dilute with water

to 1350 cc. If the solution is not clear, filter it. Keep it in a tightly stoppered bottle.

EXPERIMENTS TO ACCOMPANY CHAPTER III

13. Preparation of hydrogen. Place from 15 to 20 grams of granulated zinc in a flask arranged as shown in Fig. 119. Cover the zinc with 25 cc. of water. The thistle tube should dip below the water. Fill two or three cylinders with water for collection of the gas in the pneumatic trough. See

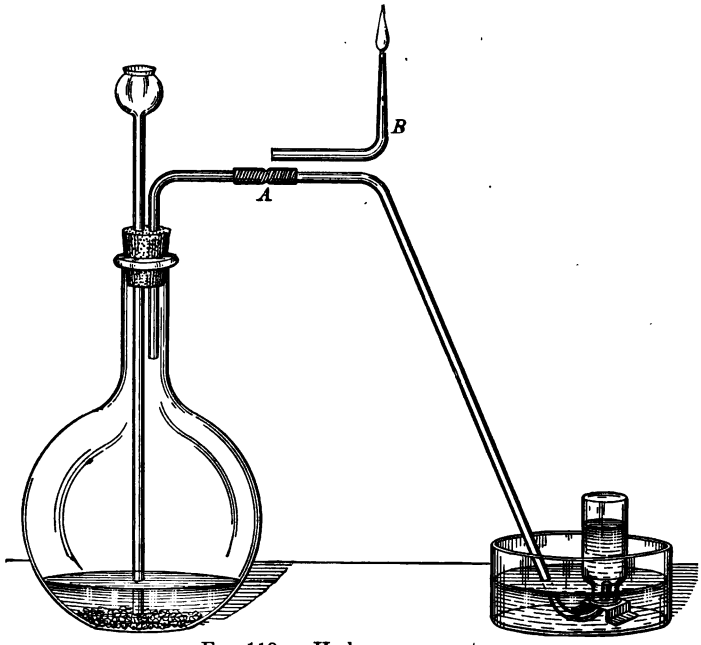


FIG. 119. — Hydrogen generator.

that the stopper is tight, for hydrogen easily escapes. When all is ready, add, through the thistle tube, about 15 cc. of concentrated hydrochloric acid. Do not apply any heat.

Do not collect any gas until the generator has run for at least two minutes. *Wrap a towel securely around the generator,* and then collect a test tube full of the gas and light it. If it burns quietly without explosion, proceed to collect a cylinder or small bottle (100 cc.) of gas. (1) Introduce a burning splinter into the gas; does it support combustion? (2) When the generator has been running for some time — five to six minutes — disconnect at *A,* attach the blowpipe tip *B,* and light the gas. *Explosions will occur if the air is not all displaced.* Hold an inverted clean dry beaker over the flame; note the production of water by the combustion. The hydrogen unites with the oxygen of the air to form water. Use your brass blowpipe as a jet when lighting the gas, and note the color of the flame. The hydrogen flame is colorless.

14. **Action of sodium on water.** In your evaporating dish place 20 cc. of water and then drop into it a piece of sodium half as large as a pea. After the action has entirely ceased, test the solution with red litmus paper. What compound is dissolved in the water?

15. **Preparation of oxygen.** Set up the apparatus as shown in Fig. 120. Fill the pneumatic trough nearly full of water and place in it the free end of the delivery tube. Weigh out 5 grams each of potassium chlorate and manganese dioxide. Mix these on a paper and place the mixture in the test tube. Have the test tube perfectly dry inside and out. Fill several cylinders with water, cover these with glass plates, and invert them on the shelf of the trough, removing the plates. With a small flame, apply heat cautiously to the test tube. The burner should be held in the hand, so that the flame may be removed from time to time; it should not be allowed to strike in one spot, otherwise the glass may crack. As soon as gas comes off freely, place the end of the delivery

tube so that the gas is collected in one of the cylinders. When the cylinder is filled, cover it with one of the glass plates while still under water. It can then be placed upright on the desk, while another cylinder is being filled. After thus filling two or three cylinders, remove the end of the delivery tube from the water, and then remove the flame. *Do not remove the flame while the delivery tube is under water.* If you

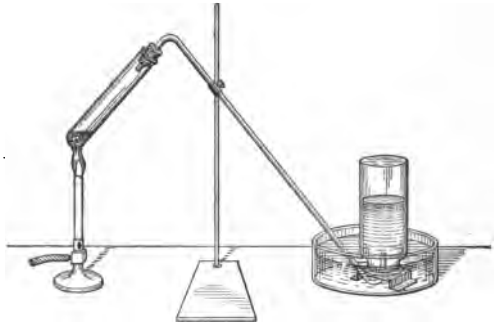


FIG. 120. — Making oxygen.

do, a vacuum will be formed and the water will be forced into the hot tube and break it. (1) Light a splinter and place it for a moment in one of the cylinders; remove it and extinguish the flame, but while still glowing, thrust it into the cylinder again. (2) Put a small piece of sulphur, no larger than a grain of wheat, into a small iron spoon. Ignite the sulphur in a flame and thrust the burning material into a cylinder of oxygen. Record the results. (3) Heat the end of a piece of fine iron wire. Dip into it flowers of sulphur, light the sulphur, and plunge the wire into a cylinder of oxygen while the sulphur is still burning. What is the result? (4) Heat a piece of charcoal in the air, and while it is still glowing drop it into a cylinder of oxygen. What is the result? Now shake the contents of the cylinder with 5 cc. of water;

then test the water with a piece of blue litmus paper. Also add clear limewater and shake the contents. Record the results and explain your observations.

16. **Ozone by electric sparks.** Work a frictional electric machine and note the odor produced in the air. Ozone has been formed by the electric discharges.

17. **Ozone.** Smell of a bottle containing yellow phosphorus in water. The odor of ozone can be detected.

EXPERIMENTS TO ACCOMPANY CHAPTER IV

18. **Carbon dioxide in the air.** Pour 10 cc. of perfectly clear limewater (calcium hydroxide) into a test tube and blow air through it, using for the purpose a clean glass tube. Observe the precipitate of calcium carbonate. Expose some clear limewater to the air in a beaker for 24 hours and observe the result. *There is a small amount of carbon dioxide in the air and this is the plant's source of carbon.*

19. **Preparation of nitrogen.** Float a thin piece of cork on a surface of water in the pneumatic trough. Place on it a piece of phosphorus half as large as a pea. Note: **Phosphorus is very dangerous. Handle it with a forceps. Keep it under water till you actually require the piece for use. Use no more than directed.** Ignite the phosphorus with the end of a hot wire, and quickly place over it a wide-mouthed bottle filled with air, being careful to keep the mouth of the bottle under water. Let the gas stand over water until the white fumes are absorbed. The white fumes are phosphorus pentoxide. Note the extent of the rise of the water in the bottle. (1) Now test the gas with a lighted splinter. Does it support combustion? (2) Place some burning sulphur on a spoon and insert it in the nitrogen. How does it behave?

20. **Preparation of nitric acid.** *Special care should be exercised by the student in the preparation of all acids;* but **with nitric acid special care must be taken,** *because if any is spilled on the hands* **it causes painful burns and wounds that heal slowly.** Arrange the apparatus as shown in Fig. 121. Carefully introduce 25 grams of sodium nitrate and then 25

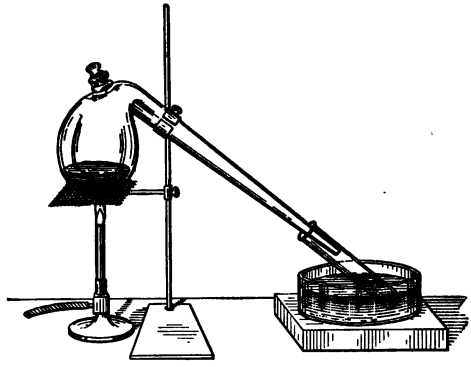


FIG. 121. — Making nitric acid.

grams of concentrated sulphuric acid into the retort, and shake very gently until all is well mixed. Heat the mixture gently and collect the nitric acid in a test tube, cooled by water. Make the following tests: (1) Remove a drop of the acid by means of a glass tube and apply it to a piece of woolen cloth or silk; observe the result. (2) Place a few drops of indigo solution in a test tube containing 5 cc. of water and then add 2 cc. of nitric acid. What happens? (3) Place a small piece of copper in the test tube containing the remainder of the acid. If no reaction takes place, add a little water. What becomes of the copper?

21. **Preparation of ammonia.** Arrange apparatus as shown in Fig. 122. Into the flask place 10 grams of ammonium chloride, 10 grams of powdered lime, and 25 cc. of water.

When the apparatus is properly connected, apply heat gently for 8 to 10 minutes. (1) Test the gas with red litmus paper. (2) Test the gas with a burning splinter. (3) Into an evaporating dish place 5 cc. of hydrochloric acid and 10 cc. of water, and pour this into a cylinder of the gas. Avoid inhaling any of the ammonia gas. The white fumes are ammonium chloride. Why did you collect the ammonia gas in an inverted cylinder? What is formed when ammonia gas is treated with water? Shake a cylinder of ammonia gas with 25 cc. of water, and test the solution with red litmus paper.

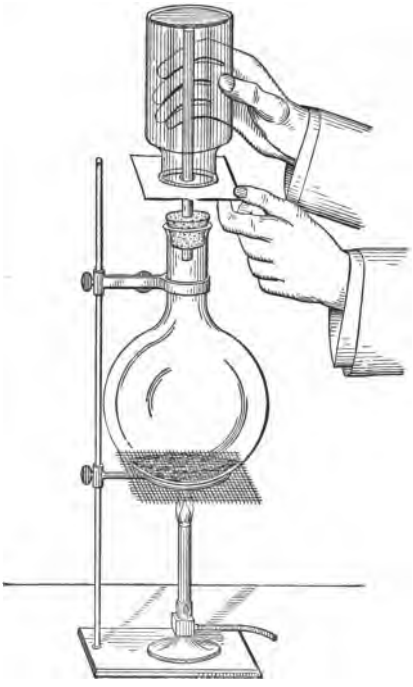


FIG. 122. — Making ammonia.

EXPERIMENTS TO ACCOMPANY CHAPTER V

22. Acids in fruits and silage. (1) Cut open a lemon and squeeze some of the juice into a beaker. Add a few drops of phenolphthalein solution and then carefully add a dilute solution of potash or baking soda until the red color just appears. Note the lowering of the acid taste as the potash is added. Lemons contain citric acid. This ex-

periment may be repeated with grapes in which the main acid is tartaric acid. From the grapes we get cream of tartar. (2) Press the juice from silage, and neutralize the acids in this in the same way. In silage the acids are mainly acetic and lactic acids.

23. **Acids and bases. Preparation of a salt.** Put 10 cc. of dilute hydrochloric acid and 10 cc. of water in a porcelain dish. Measure out 10 cc. of sodium hydroxide solution into a beaker, and add 50 cc. of water. Add this dilute sodium hydroxide to the acid, a little at a time, until the solution is neutral to litmus. Do not drop the paper into the solution, but by means of a glass rod transfer a drop of the solution to the paper. In case too much sodium hydroxide has been used, add a drop or two of the acid. *Bases or alkalies turn red litmus blue, while acids turn blue litmus red. When the solution is neutral, it has no perceptible action upon litmus paper.* Place the dish upon a sand bath and apply heat until the solution is evaporated to dryness. Avoid excessive heating. This will prevent spattering when the solution becomes concentrated. Taste some of the material in the dish. It is common salt and entirely different in taste and chemical action from either hydrochloric acid or sodium hydroxide. It is formed by the union of an acid and a base. Write the equation.

EXPERIMENTS TO ACCOMPANY CHAPTER VI

24. **Preparation of hydrochloric acid.** Into a generator, arranged as in Fig. 119, place 15 grams of common salt. Through the funnel tube add 25 cc. of concentrated sulphuric acid. Warm the mixture gently if necessary. Let some of the gas bubble through water in a receiving flask. Test the liquid with litmus paper. Taste it carefully. Take

away the water, and collect some of the gas in a bottle by displacement of air. Is the gas heavier or lighter than air? Does the gas burn? Hold a filter paper moistened with ammonia near the escaping gas. Explain, and write the equation. Invert a jar of the gas over water. What causes the water to rise? To some of the solution add a few drops of silver nitrate. A precipitate is formed which is the very insoluble silver chloride.

25. Chlorides in water. To test for chlorides, add a few drops of nitric acid to 10 cc. of water and then add 5 cc. of a 1 per cent solution of silver nitrate. The formation of a white precipitate of silver chloride indicates chlorides. Try this on well water and on service water, also on a 0.1 per cent solution of common salt.

26. Bleaching action of bleaching powder. Place 20 grams of bleaching powder in a beaker. Cover the powder with water, and then add 40 cc. of 10 per cent sulphuric acid. Hang a moist strip of red calico in the beaker by means of a wire. Now cover the beaker with a glass plate, warm gently, and let it stand. What is the result? Turkey red calico will be almost unchanged, but other cheaper calicoes will be altered.

27. Vapors of iodine. Place 1 gram of potassium iodide and an equal amount of manganese dioxide in a test tube, and then add 2 cc. of concentrated sulphuric acid. Warm gently. Note the violet color of the vapors formed. These are the vapors of iodine.

28. Vapors of bromine. Mix 1 gram of pulverized potassium bromide and 1 gram of manganese dioxide. Place this mixture in a test tube, add 2 cc. of concentrated sulphuric acid, and heat gently. Note the brown fumes of bromine which are evolved.

EXPERIMENTS TO ACCOMPANY CHAPTER VII

29. Properties of sulphur. Place 15 grams of roll sulphur, coarsely pulverized, in a test tube attached to a test tube holder, and heat slowly until it is a thin, amber-colored liquid. As the temperature increases, notice that the liquid becomes darker, until it is quite dark and so thick and viscid that it cannot be poured. Continue to heat until the material becomes slightly lighter in color and again is a liquid. Then quickly pour it into water, and when cool examine the mass and describe its properties. On standing several days it reverts to ordinary sulphur. Is sulphur soluble in water? Try to dissolve some roll sulphur in carbon bisulphide. Pour off the clear liquid on a watch glass and allow it to stand and evaporate (**no flame should be near**). Examine the crystals under a lens.

30. Sulphur dioxide. Fill a small iron spoon half full of sulphur. Ignite the sulphur and then lower it into a small cylinder containing a piece of wet, colored cloth, a piece of red litmus paper, or a red carnation. As soon as the sulphur stops burning, remove the spoon and cover the cylinder with a glass plate. What is the effect on the cloth, on the litmus paper, and on the flower? Sulphur dioxide destroys organic coloring matter and is also a good disinfectant. It kills germ life.

31. Sulphuric acid. Make the following tests with some of the sulphuric acid used in the Babcock test for milk fat. (1) Put 2 or 3 cc. in a test tube and stick a splinter of wood into it and leave it there. Then remove the splinter. Wash off the acid, and examine the stick. It has turned black. What is this black material? Also drop a piece of sugar, about a gram, into 2 cc. of the sulphuric acid. (2) Put 10 cc. of water and 1 cc. of sulphuric acid into a test tube; then add

2 to 3 cc. of barium chloride solution. Observe the result. Barium sulphate has been formed. This is a very insoluble precipitate, and *this test is a general one for sulphates.* (3) Let the barium sulphate settle, pour off the clear supernatant liquid, and try to dissolve the white salt in hot hydrochloric acid. Does it dissolve?

32. **Hydrogen sulphide.** Arrange the generator as shown in Fig. 119. The delivery tube and cork should fit tightly, and the delivery tube should pass into a solution of lead nitrate. Set the apparatus under a hood. Place 10 grams of iron sulphide in the generator, add 30 cc. of dilute hydrochloric acid, and immediately connect with the delivery tube. Note the precipitate formed. Pass the gas into a solution of common salt, then into a solution of copper sulphate, then into a solution of tartar emetic. Note the results. Note the odor of the gas. It is formed when organic matter decomposes, as in the rotting of eggs and manure.

33. **Test for phosphates.** In a 50-cc. beaker on a sand bath, dissolve half a gram of bone ash in 10 cc. of dilute nitric acid and 20 cc. of water. Filter, and to the filtrate add 5 cc. of ammonium molybdate solution and then stir the mixture. Observe the yellow precipitate, which is ammonium phosphomolybdate. *This is a general test for phosphates.* It should always be made in a faintly acid solution.

34. **Insoluble phosphates.** Dissolve one-half gram of sodium phosphate in 10 cc. of distilled water and then add 10 cc. of water containing half a gram of calcium chloride in solution. Observe the result. The precipitate is calcium phosphate. Repeat the experiment, using alum instead of calcium chloride. In this case the precipitate is aluminum phosphate.

EXPERIMENTS TO ACCOMPANY CHAPTER VIII

35. Making boric acid from borax. Boil 100 cc. of water and stir in powdered borax till no more will dissolve. Then acidify with hydrochloric acid. On cooling, the flakes of boric acid separate out. Filter off and test the solid residue as follows: (1) Dissolve some of the boric acid in 5 cc. of water contained in a porcelain evaporating dish. Add 2 drops of sulphuric acid and 10 cc. of alcohol to the solution, and then carefully apply the flame to the dish. When the alcohol begins to evaporate, ignite it with the flame. It will burn with a green flame which shows the presence of boric acid. (2) Test some of the aqueous boric acid solution with turmeric paper. Note the color produced. The result is better if the boric acid solution is acidified with hydrochloric acid and the turmeric paper is dried at 100° C. after having been moistened with this solution.

36. Insolubility of silicates. Weigh out carefully 5 grams of ordinary soil and place it in a flask and add 25 cc. of concentrated hydrochloric acid. Does it dissolve? Warm for an hour; pour off the acid through a filter paper, and wash the rest of the soil on to the filter paper. Put the funnel containing the soil in a warm place and let dry. Then reweigh the residue after tapping it out of the paper, being careful that none is lost. Has the acid dissolved very much? The soil is largely made up of silicates, which are almost insoluble materials.

37. Sodium silicate. Place 1 cc. of a sirupy solution of sodium water glass in an evaporating dish, dilute with 10 cc. of water, and add 2 to 3 cc. of concentrated hydrochloric acid. Note the gelatinous precipitate. Evaporate to dryness, and heat the dish gently over a low flame. Cool, add water, filter, and examine the solid residue. What is it?

EXPERIMENTS TO ACCOMPANY CHAPTER IX

38. Preparation of carbon. Place 2 pieces of wood and a piece of bone, each as large as a bean, in an iron crucible and cover the materials with sand. Heat the crucible until smoking ceases. Remove from the flame, let cool, and examine the charcoal and the bone black. What are the principal elements in wood? In bone? Try the same experiment with some ground wheat. Particles of carbon may be obtained from a luminous gas flame or a candle or lamp flame, by holding a piece of cold porcelain in the flame. Carbon is deposited in chimneys as soot when fuel is burned with a poor draft. When fuel is burned completely, the carbon disappears entirely as the gas, carbon dioxide.

39. Reducing power of carbon. Mix thoroughly from 2 to 3 grams of pulverized copper oxide and an equal bulk of pulverized charcoal. Place the mixture in a small test tube and apply heat. Observe the result. What is the bright red material produced? What causes the oxygen to be separated from the copper oxide?

40. Absorbing power of carbon. To 10 cc. of water in a test tube add 5 drops of cochineal solution and 10 cc. of hydrogen sulphide water. Add 3 to 5 grams of bone black or animal charcoal. Stopper the tube with a cork and shake it. Let stand 5 minutes, and then pour upon a filter. Now smell of the liquid and look at the color of the filtrate. It has lost its color, and is odorless. The charcoal absorbs both colors and odors. A half bushel of charcoal in a bag hung in the water of a cistern will gradually remove the odor of the cistern water.

41. Preparation of carbon dioxide. Arrange the apparatus as shown in Fig. 119. Put 20 grams of marble in the flask and sufficient water to cover the end of the thistle tube.

Fill 2 or 3 cylinders with water for collecting the gas. Then add slowly through the tube 30 cc. of concentrated hydrochloric acid. Allow some of the fresh gas to escape, and then collect from 2 to 3 cylinders of the gas. Remove the cylinders, and set them upright on the table. (1) While the apparatus is running pass some of the gas into a test tube of limewater, and note the result. A precipitate is first formed and gradually then redissolved. Boil the clear solution thus formed and note the reappearance of the precipitate. (2) Test some of the escaping gas with a burning splinter. (3) Pour a cylinder of the gas over a candle flame. What happens? (4) Invert another cylinder over a 25 per cent solution of sodium hydroxide. Why does the gas disappear?

Agricultural students may also perform similar experiments, using other calcium-containing substances, such as caustic lime, limestone, floats, and gypsum, to see if these contain carbonic acid, as does marble.

42. **Making hard soap.** In a beaker dissolve 15 grams of caustic soda in 120 cc. of water. Pour half of the water into a porcelain dish of about 500 cc. capacity, add 50 cc. of water and 50 grams of tallow. Boil the solution for 45 minutes, carefully replacing from time to time the water lost by evaporation. Then add the remainder of the caustic soda and boil for at least an hour more. The volume of the liquid may finally be allowed to decrease about one-third. Add 20 grams of common salt, boil for a few minutes, and allow to cool. The soap rises to the top, and the glycerine, excess of lye, and salt remain in solution. Try to make soap similarly, using gas engine cylinder oil instead of tallow.

43. **Alcoholic fermentation.** Weigh 10 grams of flour into a 500-cc. bottle, add 50 cc. of water and a small piece

of a yeast cake. By means of a delivery tube connect the flask with a test tube containing enough clear limewater to cover the end of the tube (see Fig. 123). Cover the limewater with one-quarter of an inch of kerosene. Why? Place the bottle on a warm sand bath (not over 85° F.) for half an hour. Why? Observe the bubbles of gas given off and the precipitate that is formed in the limewater. Carbon dioxide and alcohol are formed. This is what occurs in bread making.

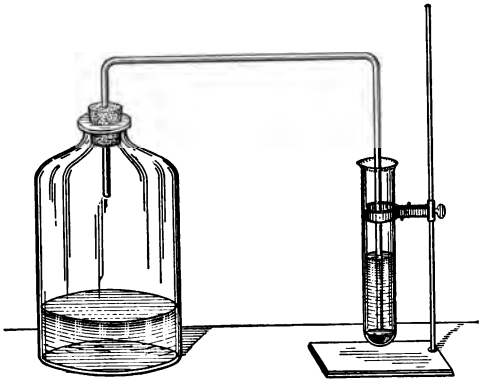


FIG. 123. — Alcoholic fermentation.

44. **Removal of grease spots from clothing.** Spots due to fats such as butter, tallow, olive oil, etc., can be removed with soap or with gasoline or benzine; also by placing the fabric between blotting paper and applying a warm iron. Try the solubility of fats in gasoline or benzine, also the removal of a grease spot by the hot iron method. Grease may also be removed by ammonia water or even a solution of borax. In the case of spots due to mineral oils, the spot should first be treated, or softened with some oil and then treated with gasoline. Why? **Gasoline is inflammable. Do not have flames around when you are using it.**

45. **Washing soda.** This is sodium carbonate and is sometimes used as a filler for soap. It is the main constituent of washing powders. To test such a powder for sodium carbonate, put 5 grams in a test tube and add a few drops

of hydrochloric acid. If there is a brisk evolution of gas, it indicates the presence of a carbonate.

46. Odors and appearance of some important organic compounds. The instructor should have the student examine with the eye and nose the following materials: carbolic acid, ether, banana oil, carbon tetrachloride, chloroform, etc.

47. Test for nitrogenous organic compounds. Mix a gram of dry clover hay, peas, or meat with enough soda-lime to half fill a test tube. Connect the test tube with a delivery tube as in Fig. 120, but let the end of the tube lead into another test tube containing water. Apply heat for from 5 to 10 minutes. Test the water with red litmus paper. It should turn blue. Soda-lime decomposes protein material and liberates nitrogen as ammonia. The carbon and oxygen of the protein unite and form carbon dioxide and water; the carbon dioxide forms carbonates with the soda-lime.

48. Test for proteins. Make about a 10 per cent caustic potash solution by dissolving 5 grams of caustic potash in 45 cc. of water. In another glass dissolve a piece of copper sulphate (bluestone) the size of a pea in 50 cc. of water. Pour about one-fifth of the white of an egg into 50 cc. of water. Place 10 cc. of this egg solution in a test tube, add 5 cc. of the potash solution, put the thumb over the tube, and shake until all is well mixed. Now add a few drops of the copper sulphate solution and shake again. At first the only color will be a greenish blue, but on standing this will change to a deep violet color. *This is the test for proteins.* Test wheat, corn, bran, and meat by this means. These materials should be crushed or ground, and warmed for a short time in a little of the caustic potash solution and then filtered before adding the copper sulphate solution.

49. **Preparation of a protein.** Put about 25 grams of flour in a porcelain dish, add 12 to 15 cc. of water, and mix into a stiff dough. Do not have too much water present. Let stand one-half hour, remove the ball of dough, and, holding it in the fingers, add water little by little and so wash out the starch by further manipulation with the fingers. Continue the washing until the liquid runs clear, and then place the ball of gluten in water for one-half hour. Remove it from the water, freeing it from water as much as possible, and then weigh it. Spread it out in a thin cake, and then let it dry in the oven or in a warm place, and finally reweigh it. *Calculate the per cent of gluten in the flour.* The gluten consists of the two principal proteins of the wheat.

50. **Fats and oils.** Crush 5 grams of flax or hemp seed on a piece of white paper. A grease spot will appear on the latter. Should the test fail to show oil, place the crushed seeds and the paper in a warm place on a tin plate. The higher temperature will melt the oil out, which will then be shown by the grease spot on the paper.

51. **Fats and oils.** Crush four or five kernels of corn. Place the powder in a flask, and pour over it 25 cc. of gasoline. Shake it at intervals for 15 minutes. Keep the flask loosely stoppered and **away from flames.** Then pour the gasoline off, and through a filter paper, catching the filtrate in a beaker. Allow the gasoline to evaporate slowly. A residue is left in the beaker. Test it between the fingers for its greasy feeling. Gasoline and benzine dissolve fats. Grease spots on cloth are removed by washing with gasoline or benzine.

52. **Preparation of starch.** Make a pulp of two clean potatoes. Tie the pulp in a cloth and squeeze the juice into a beaker filled with water, occasionally dipping the bag into water. Allow the beaker to stand for 20 minutes, or

until the starch has settled. Now pour off the water, and if the starch is not clean, wash it by adding more water and again allowing it to settle. Finally pour off the water. Leave the beaker in the desk until the starch is dry, and then save the starch for the following tests.

53. **Test for starch.** Taste some of the starch. It is practically tasteless. Place one-half gram in a test tube about half full of water. Shake the test tube, then boil, and filter. To the filtrate add a few drops of iodine solution. (This is made by dissolving 5 grams of potassium iodide in 50 cc. of water and then adding 1 gram of iodine to the solution.) A deep blue color appears throughout the solution. Treat some of the starch with cold water, but do not filter; add some of the iodine solution. On standing, the starch, settling out at the bottom, will be blue. In the first case the starch was in solution, in the second it was not. Preparations of starch for the iodine test may also be made from corn, wheat, oats, or beans.

54. **Test for sugar (glucose).** Grind 25 grams of dried raisins in 50 cc. of water. Filter the solution into a beaker. Into a test tube place 10 cc. of **Fehling's solution.** This is made by dissolving 6.2 grams of copper sulphate, 3.5 grams of Rochelle salts (sodium potassium tartrate), and 2 grams of potassium hydroxide in 100 cc. of water. Place about 10 cc. of the Fehling's solution in a test tube and add from 5 to 10 cc. of the raisin solution. Heat over a flame and let the liquid boil a few seconds. A reddish brown precipitate forms, which on standing deposits on the bottom of the test tube. This precipitate is the red oxide of copper, Cu_2O, and shows the presence of glucose in the solution. Test the water extract from corn or wheat for glucose; also the extract of an apple, a sample of honey, and one of "Karo sirup."

EXPERIMENTS TO ACCOMPANY CHAPTER X

55. Preparation of potassium hydroxide. Dissolve 5 grams of potassium carbonate or wood ashes in 15 cc. of water in an evaporating dish. Add a mixture of 3 grams of barium hydroxide (also called barium hydrate) and 10 cc. of water. Heat on a sand bath for 5 minutes. Filter off the solution and observe the precipitate. Evaporate some of the filtrate to dryness by heating on the sand bath. This is potassium hydroxide. *Keep it for experiment 56.* Wood ashes contain potassium carbonate and are often the source of the "lye" used in soap making on the farm.

56. Test for potassium. Dip a *clean* platinum wire into a solution of the residue left from the last experiment. Hold it in a non-luminous Bunsen flame and observe the flame through a blue glass. A violet color shows the presence of potassium. If no platinum wire is at hand, use a clean strip of asbestos paper. Hold a crystal of potassium chlorate in the flame with the forceps.

57. Test for sodium. Dip a clean platinum wire into a solution of common salt; hold it in the Bunsen flame and observe the bright yellow color imparted to the flame. This is a test for sodium. Try the same test by holding a piece of rock salt in the flame with a forceps.

58. Precipitation of barium sulphate. To a solution of barium chloride add a drop of sulphuric acid. Note the heavy white precipitate formed, which is the insoluble barium sulphate. To a solution of copper sulphate add a few drops of barium chloride. Here again the insoluble barium sulphate is formed.

59. Tests for the alkaline earth metals. Using a clean forceps, hold a piece of a salt of strontium in the flame and note the color produced. Perform the same experiment,

holding a piece of a calcium salt in the flame; also repeat the experiment with a piece of a barium salt.

60. Testing the quality of lime. Place about 40 grams of burned lime in an evaporating dish, and moisten it well with water, warming to about 35° C. to hasten the action, if necessary. Note the reaction and observe how the mass warms up. Good, fresh lime readily undergoes the slaking process. Place some of the slaked lime in a bottle; add 100 cc. of distilled water; shake vigorously, and allow to stand for four or five hours. Then filter some of the solution and test it by blowing air through it, using a glass tube. Place about one-half gram of the slaked lime in a test tube, add 10 cc. of water and then a few drops of hydrochloric acid. When the action ceases, add more acid, a little at a time, and warm. What has not dissolved consists of silica and clay. Lime of high purity contains less than 10 per cent of acid-insoluble material.

61. Plaster of Paris. In a test tube carefully heat 10 grams of powdered gypsum to 115° C. Regulate the temperature by means of a thermometer so that it reaches no higher than 115° C. Why? Allow it to cool. Mix with a small quantity of water and note that in a few minutes the mass becomes hard. Explain.

62. Burning magnesium. Hold a piece of magnesium ribbon about an inch long in a forceps and apply a lighted match. The ribbon burns with a brilliant white flame. Examine the white product formed. What is it? Flash-light powder consists of about 5 parts of powdered magnesium to 9 parts of potassium chlorate. It should be used with care.

EXPERIMENTS TO ACCOMPANY CHAPTER XI

63. Testing for alum. Add a few drops of an alum solution to a test tube containing 5 cc. of water and then add a few drops of logwood extract and 2 cc. of ammonium carbonate solution. Observe the result. Mix about 2 grams of flour in a dish with water containing a few drops of alum solution. Add a few drops of logwood extract and the same amount of ammonium carbonate solution. Mix well and observe the result. Repeat the test, using a baking powder, testing for the presence or absence of alum. *In the presence of alum a blue color is always obtained with tincture of logwood and ammonium carbonate solution.*

64. Compounds of alum and protein. To 5 cc. of a solution of egg white add a few drops of an alum solution and observe the result. A heavy precipitate will form. This is an insoluble compound of protein and alum. To a solution of aluminum chloride add ammonia water. The same result is noticed, but in this case aluminum hydroxide is formed.

65. Copper in a silver coin. Under a hood, dissolve about one-fourth of a dime in dilute nitric acid. It is best to hammer out the coin and then cut it. Account for the bluish color of the solution. To the solution add a solution of hydrochloric acid until no further precipitate is formed. The white precipitate is silver chloride. Filter and evaporate the blue solution of copper nitrate in an evaporating dish, finally igniting to redness over a flame. The black residue is the oxide of copper. Expose the silver chloride to the light. What happens?

66. Copper and its reactions. (1) Dissolve 5 grams of copper sulphate in 100 cc. of water. Place a bright iron nail in the solution and note the result. (2) To 5 grams of

Rochelle salts add 10 cc. of water, and 5 cc. of concentrated caustic soda solution, after all the salt has dissolved. Now add 5 cc. of the copper sulphate solution and then treat the resulting solution with 1 cc. of glucose ("Karo sirup") and boil. A reddish brown precipitate of cuprous oxide, Cu_2O, is formed. This is the test for sugar. (3) To 10 cc. of the copper sulphate solution add a few drops of potassium ferrocyanide. A reddish brown precipitate is formed. This is a delicate test for copper, and can be used for testing Bordeaux mixture for excess of copper sulphate in solution. In Bordeaux preparation, if this test shows the brown precipitate, more lime must be added; see experiment 151.

67. **Copper in brass.** Under the hood, treat about 0.2 gram of brass with 5 cc. of nitric acid in a test tube and note the blue color of the solution. This color is due to the formation of copper nitrate.

68. **An antidote for mercury poisoning.** To 20 cc. of water add 5 or 6 drops of the white of an egg. Stir thoroughly and then add to this solution a few cubic centimeters of a dilute solution of mercuric chloride (corrosive sublimate), 1 in 2000. The precipitate is a compound of the protein and mercury. *In case of mercury poisoning the whites of eggs and milk are a good antidote.*

69. **Preparation of mercury.** To 5 cc. of a solution of stannous chloride add a few drops of a solution of mercuric chloride. The gray precipitate that separates is metallic mercury. Note the physical properties of some mercury contained in a bottle on the reagent shelf.

70. **Lead and its salts.** Dissolve 5 grams of metallic lead in warm, dilute nitric acid. Evaporate off nearly all of the excess of acid and allow the liquid to stand and cool. Pour off the remaining liquid from the crystals and dissolve these in 50 cc. of water. Test separate portions of 5 cc. each of this

solution with (*a*) hydrogen sulphide; (*b*) dilute sulphuric acid; and (*c*) dilute hydrochloric acid. Precipitates of different colors are formed. Lead sulphide is black; lead sulphate and chloride are white.

71. **Action of water on lead.** Put a gram of clean bright lead shavings into a test tube containing 10 cc. of distilled water. After 24 hours decant the clear liquid into a second tube. Slightly acidify with hydrochloric acid and then add 5 to 10 cc. of fresh hydrogen sulphide water. A black or brown coloration indicates lead in solution. Soft water, especially peaty water, attacks and dissolves lead more than does hard water. In contact with hard water the lead becomes coated with the insoluble lead sulphate and carbonate of lime which protect the metallic surface from further attack.

72. **Preparation of lead chromate.** To 5 cc. of a solution of potassium dichromate add 3 cc. of lead acetate (sugar of lead) solution. The beautiful yellow precipitate formed is chrome yellow, or lead chromate, which is used as a pigment in paints.

73. **Making formalin.** Treat 10 to 20 cc. of wood alcohol with one-half gram of potassium permanganate and 1 to 2 drops of strong sulphuric acid. Heat the beaker gently. Note the odor produced. The penetrating smell is due to the production of formaldehyde, formed from the wood alcohol by oxidation of the latter by the permanganate.

74. **Cobalt bead.** In a loop of platinum wire, made by bending the wire around the sharp end of a pencil, place some borax and then fuse it in the flame. Note the color of the bead. Now dip the bead into a dilute solution of cobalt nitrate. Then put the loop with the bead into the flame again and allow the bead to melt. Note the beautiful blue color of the bead on cooling.

75. Preparation of iron sulphate. Place 25 grams of iron nails in a 500-cc. flask, add about 25 cc. of dilute sulphuric acid (1 part acid to 5 parts water), warm gently until no more gas comes off. Pour the clear solution into a clean beaker. Add 2 to 3 cc. of dilute sulphuric acid to the solution, boil the liquid down to one-half of its volume, and allow it to cool. Crystals of iron sulphate will separate out. The liquid should be poured off and the crystals allowed to drain. This is the material used in the destruction of weeds and is commonly prepared at steel plate mills. Other names for this salt are *copperas, green vitriol, and ferrous sulphate.*

76. Change of ferrous to ferric iron. Dissolve one-half gram of ferrous sulphate in 20 cc. of water. Divide the liquid into two equal portions. To the first portion add a few drops of ammonia water until a precipitate is obtained. To the other portion add five drops of concentrated nitric acid. Heat this latter portion to boiling; cool, and then add ammonia water until a precipitate forms. Compare the two precipitates. The nitric acid oxidizes the iron, which then is in a substance containing more oxygen. The copperas is ferrous sulphate; by oxidizing it with nitric acid it was changed to ferric sulphate.

77. Compounds of tannin and iron (ink). (1) Dissolve 1 gram of tannic acid in 25 cc. of hot water. Dip a piece of cotton cloth (about 3 by 10 cm.) into the solution. Dry the cloth, and then dip it into a solution of ferrous sulphate (1 gram in 25 cc. of water). After the cloth has dried, see if the color can be removed by washing. (2) Add 5 cc. of ferrous sulphate solution to 5 cc. of tannic acid solution. Observe the result. Does it look like ink?

78. Making a blue print. Place a leaf or a drawing made on thin paper on the glass of a photographer's printing frame and then cover it with a piece of blue printing paper (the

sensitive side next to the leaf). This paper can be secured at most drug stores. The operation of placing the object and paper together should be done in a darkened room. Now expose the printing frame to direct sunlight for about 3 to 5 minutes. Several trials must be made to determine the time of exposure necessary for a successful print. Remove the paper from the frame. Wash with water and spread the print on a cloth to dry.

EXPERIMENTS TO ACCOMPANY CHAPTER XII

79. Drying of linseed oil. With a brush spread a layer of linseed oil (very thin) on a well-varnished ruler and allow it to stand several days. Now examine it. Has it set to a hard film? Crush 10 grams of flaxseed in a mortar. Extract the material with gasoline in a flask. Pour off the gasoline, and out of doors, or under a hood, allow the gasoline to evaporate off. Paint the ruler with the oil that is left. Try a film of corn oil on a ruler. How does it act?

80. Solubility of oils. Try to dissolve some linseed oil in gasoline; also in carbon bisulphide. Try the same experiment, using corn oil, also butter fat, instead of linseed oil.

EXPERIMENTS TO ACCOMPANY CHAPTER XIII

81. Odor of burning wool and cotton. Burn small pieces of wool, silk, and horse's hoof and note the odor in each case. Try the same experiment with cotton. The former materials are nitrogenous and give rise to ammonia and other odoriferous nitrogenous compounds, while cotton does not. Try the solubility of the various materials mentioned above in ammonia water. Wool, silk, and horn will dissolve, while cotton will not.

82. **Bluing.** Indigo was very generally used for bluing in the laundry. Another material is now commonly employed; namely, *Prussian blue.* This is ferric ferrocyanide, a complex cyanide of iron. Because of the ease with which it decomposes with alkalies, there is danger that iron rust may be deposited on the goods if this form of bluing is used. To 10 cc. of a 10 per cent solution of ferric chloride, add 5 cc. of a 10 per cent solution of potassium ferrocyanide, and about 5 drops of hydrochloric acid. The blue precipitate is Prussian blue. Now add an excess of caustic alkali and heat to boiling. Note the reddish brown precipitate of ferric hydroxide. This is what causes the iron rust stains.

83. **Dyeing cloth.** Dip strips of white woolen, silk, and cotton cloth into a solution of fuchsine. After immersion in the solution remove the strips and allow them to drain. When they are nearly dry they may be washed in water to test the fastness of the color on the various fabrics.

84. **Gelatine in sole leather.** A piece of sole leather, half as large as the hand, should be boiled in water for an hour and then allowed to cool. Does the solution gelatinize or become viscid and sticky? Sole leather has been treated in such a way that it contains no gelatine. Try the same experiment, using a piece of fresh cowhide. Note the result.

85. **Solubility of rubber.** Place about 0.5 gram of rubber in a test tube, add 5 cc. of carbon bisulphide, heat gently in a hot water bath, and shake the contents of the tube. What happens? Notice how the rubber swells up and slowly dissolves. **No flame should be near.**

86. **Solubility of asphalt.** Dissolve some asphalt paint in turpentine and apply it to a piece of unpainted wood. Set the wood aside to dry and note the stain left behind.

87. **Removal of iron rust from cloth.** Stretch the stained cloth over a dish containing hot water. Then as the steam

rises and the cloth becomes moist, drop a few drops of dilute (10 per cent) hydrochloric acid or, better, 10 per cent oxalic acid upon the rust spot, preferably with a medicine dropper. When oxalic is used the fabric may be immersed directly in the acid. After a moment, lower the fabric into the water. If the spot is not removed, repeat the operation. Then rinse in clear water and finally in dilute ammonia water to neutralize any remaining acid. Then rinse again. Iron rust stains may often be completely removed from fabrics by means of lemon juice.

88. **Removal of paint stains.** These can be removed by the use of turpentine. Spot some cloth with paint and, after it is thoroughly dried, remove the stain with turpentine.

89. **Removal of ink stains.** A great many ink stains can be removed with sour milk or with some dilute acid such as oxalic or citric acid. If the ink is an iron tannate, the methods of removing iron rust will be applicable. Stain some cloth with ink, and then remove the spot with the reagents mentioned. These should be used in dilute solution (5 to 10 per cent); and after application and removal of the stain the fabric should be well washed with water. Long contact with the acid may weaken the fabric.

90. **Removal of fruit stains.** Fruit stains may be removed from many fabrics by washing in water containing a little borax or ammonia. Always try this treatment first, as it will not injure the colors. If the above does not remove the stain, try a very dilute solution of chloride of lime containing a few drops of dilute acetic acid (vinegar). When the stains have been removed, wash the fabric thoroughly with clean water. In the case of wool or silk, treatment with soap (care) and water is the only way. If not effective, the fabric must be re-dyed. Try these tests, and also the action of ammonia on wool and silk.

91. Coffee and tea stains. These are removed in the same manner as the fruit stains. They are due to the coloring matter in the liquids. *Grass stains* owe their color to chlorophyll. This is soluble in ammonia water or in alcohol, and can therefore be removed from fabrics by these reagents. Apply these reagents to such stains. Also try leaving the stained material in the sunlight. What is the result?

92. Blood stains. These are due to the red pigment of the blood called hæmoglobin. This coloring matter is easily decomposed by acids. It is also soluble in warm water, which should be tried for the removal of the stain, before using the acid treatment.

EXPERIMENTS TO ACCOMPANY CHAPTER XIV

93. Nitrogen in soil. Mix 5 grams of soil and an equal bulk of soda-lime in a mortar and transfer the mixture to a strong test tube. Connect the tube with a delivery tube as in Fig. 120, but have the end of the tube lead into another test tube containing distilled water. Heat the first test tube cautiously with a burner for five to ten minutes. Test the water with litmus paper and note the reaction. The soda-lime, aided by heat, decomposes the organic matter of the soil and forms water, carbon dioxide, and ammonia. The latter contains the nitrogen and has passed over into the water.

94. Phosphorus in soils. Place 10 grams of soil in a flask, add 25 cc. strong hydrochloric acid, and warm on a water bath for half an hour. Pour off the acid, cool, and add ammonia gradually, shaking after each addition, and continue this until the solution turns red litmus blue. Then add nitric acid drop by drop until the solution is slightly acid,

or until the blue litmus becomes red. Then add 10 cc. of ammonium molybdate solution. A yellow color or precipitate after a few minutes shows the presence of phosphorus. It is best to keep the solution at 65° C. to aid the precipitation of the *ammonium phosphomolybdate.*

95. Reaction of soils. For this experiment use (1) peat, (2) a mildly alkaline soil, and (3) a clay soil. Bring in contact with each soil, moistened with distilled water, pieces of sensitive red and blue litmus paper. Note any changes in color of the litmus paper and record what results you find. In a similar way test the soils from the fields of your own farm.

96. Correcting the acidity. When you find an acid soil, — that is one which changes blue litmus to red, — take a wash basin full of it and add 15 to 25 grams of burned lime or ground limestone. Stir the mixture thoroughly and test it again. Try the same experiment, using wood ashes instead of lime. If the soil is still acid, add enough lime until it shows a blue color with litmus. An alkaline or "sweet soil" is the normal one for most agricultural plants.

97. Weathering of limestone. Place about 25 grams each of fresh limestone, rotten limestone, and residual limestone soil in separate beakers. Add 100 cc. of 5 per cent hydrochloric acid to each, and allow them to stand for one hour. Filter off the insoluble material on filter papers already dried and of known weight. Wash the residues with water. Dry them in the oven at 100° C. and weigh. Calculate the per cent of insoluble matter in each case and explain the results.

98. Puddling of clay. In each of two basins place a pound of dry clay soil. To one of these add water until the soil is slightly moist, but do not work or stir it. To the other add enough water to make the clay sticky; mix thoroughly

with a stick and place the mixture in the sun to dry. Note the results.

99. Effect of soluble salts on soils. To 20 pounds of soil in a box, add 25 grams of sodium nitrate and mix thoroughly. To another box of the same soil add 2 grams of the same salt and mix. Water both, and set them in a proper place for growth. Plant with the same number of seeds (25) and after the plants are well started thin to the same number (12). Note the results. Too much of soluble salts may be harmful. An acre of soil 6 inches deep weighs about 2,000,000 pounds; estimate the application per acre on the above amounts of nitrate used for 20 pounds of soil. Does this suggest why soluble fertilizers must not be applied in large amounts?

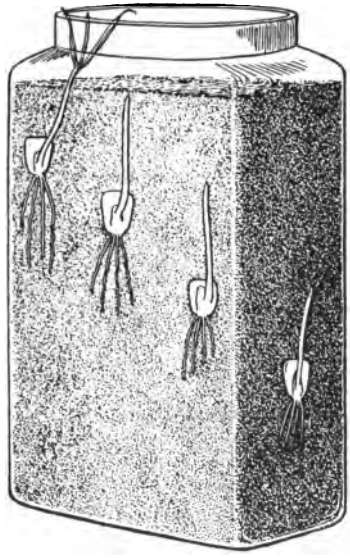


FIG. 124. — Effect of deep planting on germination of seed.

100. Effect of deep planting on the germination of seeds. Place 2 to 3 inches of moist soil on the bottom of a glass jar. Plant 4 to 5 peas near the wall so that they may be observed. Now cover with 2 inches more of soil and plant some more seed. Repeat this (Fig. 124), planting the seeds only an inch deep. Keep the soil moist and warm. Record what happens.

101. Capillary action of water on soil. Firmly tie a piece of linen cloth over the ends of several glass tubes, one-half to three-fourths inches in diameter. Lamp chimneys are suitable. Fill one tube

with sandy soil, another with clay soil, and a third with a loam. Compact the soils after each addition by gently jarring. Then immerse the lower ends of the tubes in a basin of water. Support them upright and allow them to stand. For one week, measure each day the height to which the water rises. What does this show?

102. **Sedimentation of clay.** In each of three separate cylinders, or beakers, place 200 cc. of turbid water made by shaking up some clay. To beaker No. 1 add one-half gram of burned lime and stir; to No. 2 add one gram of burned lime and stir; the third beaker is used for purposes of comparison, and no lime is added to it. After 24 hours observe the beakers, and note the influence of the lime in throwing down the clay and clarifying the liquid. Liming clay soils makes them work better.

103. **Effect of color on soil temperature.** Fill two boxes, each one foot square and five inches deep, with soil and gently compact this. Cover one with lampblack and the other with gypsum or chalk. Insert thermometers in the soil to a depth of 1.5 inches and expose the boxes equally to the rays of the sun, recording the reading of the thermometers every ten minutes for two hours. Arrange the results in a table. What is the conclusion?

104. **Effect of drainage on soil temperature.** Fill two boxes as in the above experiment, but have one box lined with tin or galvanized iron or oiled cloth. Add water to the soil in the box not lined until it begins to drain. Add the same amount to the other. Let stand until the water in the unlined box has largely drained away. Take temperature readings as in experiment 103.

EXPERIMENTS TO ACCOMPANY CHAPTER XV

105. Action of fertilizers. Into each of two boxes, about one foot square and five inches deep, place 20 pounds of ordinary dry soil. Before putting the soil into one of the boxes, add to it the following materials, mixing them thoroughly with the soil: 5 grams of calcium acid phosphate, 3 grams of potassium nitrate, and 3 grams of potassium sulphate. Water the soils, and when they have been properly worked, plant in each box 25 rape seeds and place the boxes in a greenhouse or window. When the plants are up, thin them down to 16 and let them grow. Note the difference in growth from week to week. Try the same experiment, using sand as the soil, and note the results. What elements needed by the plants may be lacking in the sand, and are not supplied by the fertilizer you have used? *Action on acid soil:* try the same experiment, using an acid soil and clover seed instead of rape seed. To one of the boxes add 10 grams of finely ground limestone. Why add the limestone?

106. Phosphorus in bones. (1) To about one gram of bone ash in 15 cc. of water add 3 to 5 cc. of nitric acid, shake the mixture and filter. To the warm filtrate add 5 to 10 cc. of ammonium molybdate, and warm. What is the yellow precipitate? What does it indicate? (2) In a test tube heat one gram of bone ash with 20 cc. of distilled water and filter. To the warm filtrate add 5 cc. of ammonium molybdate solution. Note the result, and compare it with that obtained in the previous experiment (1) in which nitric acid was first added to the ash. Try the same experiment with acid phosphate; also with floats.

107. Solubility of potash fertilizers. Place 1 gram of kainit in a test tube, add 100 cc. of water, and shake the mix-

ture. Does the salt dissolve? Try the same experiment with saltpeter (potassium nitrate); also with muriate of potash. These salts are all readily soluble potash fertilizers.

108. **Making acid phosphate.** In a cigar box mix 100 grams of bone meal and 100 grams of commercial sulphuric acid. Add the acid slowly, stirring constantly with an iron rod. After all has been added, allow the mixture to stand for three days. Then pulverize and examine it. Test 1 gram of it for soluble phosphates as in experiment 106. Also test the solubility of floats in water.

109. **Differences in volatility of ammonium salts.** In a test tube place 2 grams of ammonium carbonate. Note the odor. Apply heat gently and observe the result. The ammonium carbonate volatilizes readily and may again deposit on the cold walls of the test tube. Try the same experiment with ammonium sulphate. This is much less volatile. In poorly ventilated barns, deposits of ammonium carbonate are frequently found, especially in horse stables. Gypsum, *i.e.* land plaster, when applied to the manure converts some of the ammonium carbonate to ammonium sulphate and thus saves it.

110. **Solubility of nitrogenous fertilizers.** Place 10 grams each of sodium nitrate, ammonium sulphate, and dried blood on filter papers in separate funnels. Pour over each sample small portions of water, and continue to leach the samples with water until 100 cc. of filtrate have been collected in each case. Which of these fertilizers are insoluble in water?

EXPERIMENTS TO ACCOMPANY CHAPTER XVI

111. **Losses from manure by leaching.** Take one pound of fresh manure, place it in a large funnel having a good wad of cotton in the bottom, and leach the manure by pouring on water until one or two liters of filtrate have been collected. In a box one foot square and five inches deep, containing 15 to 20 pounds of rather infertile soil, plant 25 rape seeds. In a similar box containing the same quantity of soil, but to which the dung washings will be added as needed, also plant 25 rape seeds. After the seeds have germinated, thin down to 16 plants. Keep both boxes properly watered in a greenhouse or near a window and note the differences in growth. This experiment gives an idea of the great losses in plant food that may occur by the leaching of manures.

112. **Loss of ammonia from manure.** Place some fresh horse manure in a glass bottle and stopper it. Let it stand in a warm place for 24 hours, then open the bottle and note the odor. The penetrating odor is due to ammonia. Repeat the above experiment, but over the partly compressed manure place a layer of moist dirt one inch thick. Is the odor of ammonia so noticeable? Moist earth is a good absorbent for ammonia.

113. **Plant food in urine.** Test some urine (human or animal) for potassium by the flame test as in experiment 56, also for sodium as in experiment 57. In a test tube place about 8 cc. of urine; add a few drops of nitric acid and 5 cc. of ammonium molybdate solution. Warm to 65° C., but do not boil. A yellow coloration or precipitate shows the presence of phosphorus. Why? There is very little phosphorus in the urine of our domestic animals. Most of the phosphates leave the body in the solid excreta.

EXPERIMENTS TO ACCOMPANY CHAPTER XVII

114. Oxygen given off by leaves. Take a small bunch of young, green, active shoots of any plant, and place it in the apparatus as shown in Fig. 125. Set the jar in the brightest light available, and repeat the experiment with another jar kept in the dark. Arrange a third similar experiment in the light, with water that has been boiled and then quickly cooled. As the light falls on the leaves in the unboiled water, little bubbles of gas will begin to appear at the tips of the leaves. As they accumulate they will ascend into the test tube. No such bubbles will appear from the leaves in the dark, or from those in the light that are immersed in boiled water. Test the gas in the test tube for oxygen with a glowing splinter. Explain the results of these three experiments.

FIG. 125.— Oxygen is given off by leaves.

115. Carbon dioxide taken up by the leaves. Arrange an apparatus as shown in Fig. 126. Fill the glass tube, which is from 5 to 6 feet long, with freshly gathered leaves of grass. Use only young, active leaves. Set up the tube horizontally in the brightest daylight available, and with the aspirator slowly draw air first through the tube and then through the wash bottle containing some barium hydroxide solution.

Barium hydroxide is very sensitive to carbon dioxide, becoming milky with a small quantity of the gas. If the light is good there will be no evidence of carbon dioxide in the air that has been drawn over the leaves. Now detach the tube of leaves and draw the ordinary air through the tube. Milk-

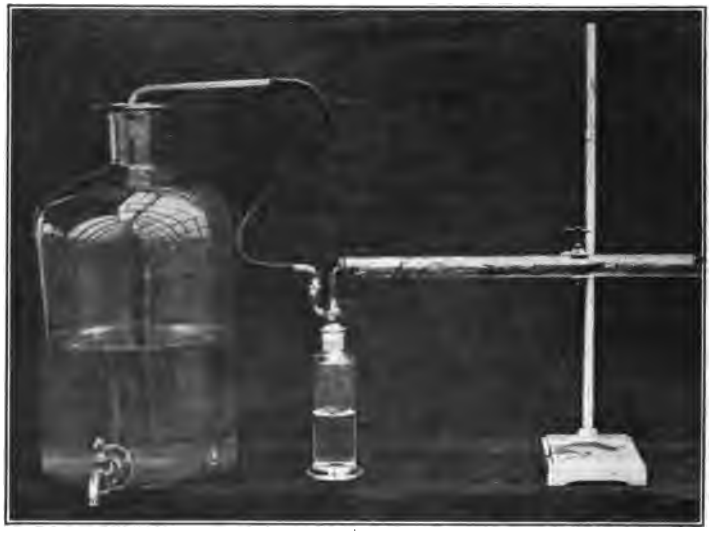


FIG. 126. — Green leaves take carbon dioxide from the air in the sunlight.

iness will soon appear. This experiment shows that green leaves take carbon dioxide from the air in the sunlight.

116. Germination of seeds. Take about 20 conveniently large seeds, such as beans or corn, and soak them in water for a few hours. Then plant them in moist sawdust, covering the seeds about half an inch deep. Keep them in a warm place for a day or two in order to start germination. Examine the seeds from time to time and note their progress. How long does it take for the seeds to grow a radicle an inch long? When does the plumule appear?

117. Influence of poisons and mutilation on germination. Soak half a dozen beans for a few hours and then mutilate some of them in various ways before planting them in moist sawdust. Cut into the embryo of one of the beans; cut bits off from the embryo of another; poison a third by touching it with a trace of carbolic acid; treat another with a drop of boiling water; etc. Then set them out in the moist sawdust to germinate. The seed will not grow if the embryo is cut or killed. The healthy embryo itself may grow when separated from the rest of the seed, provided the embryo is kept moist with a proper food solution containing sugars, proteins, and salts.

118. The germinating seed needs air. (1) Put some mustard seed into a bottle. Fill it completely with water. Then stopper it and let it stand in a warm place. The seeds may start to sprout because of the dissolved air in the water, but they will soon stop growing. (2) Grease the stopper of a wide-mouthed pint bottle. Put into the bottle one-half ounce of mustard seed and enough water to moisten the seeds thoroughly. Stopper the bottle and set it in a warm place in the dark. The seeds will germinate, for there is some air in the bottle; but they will soon stop growing further for lack of air. Open the bottle and at once insert a lighted taper. The flame is extinguished, indicating that the oxygen in the bottle has been used up. (3) Decant some of the air from the bottle just used into a clean bottle and shake it with clear limewater. The solution becomes milky, indicating the presence of carbon dioxide. Explain the results.

119. Carbon in plants. Take some dry plant tissue from growing plants, like beans or wheat, and place it in a porcelain dish. Heat it strongly over a Bunsen flame, covering the greater part of the dish with a porcelain or iron cover. Thick gases will come off, and these will take fire if allowed

to come into contact with the flame. After a time, extinguish the flame, by completely covering the basin. Withdraw the burner, and allow the dish and its contents to cool. It will be seen that the interior of the dish is covered with a black soot which is carbon, set free from the plant tissue.

120. **Nitrogen in plants.** The important element nitrogen, which is contained in the plant in combination with other elements, is lost during combustion and therefore is not in the ash of a plant. To show its presence in the plant mix a gram or two of dry plant material in a test tube with soda-lime, and then heat the whole directly in a flame; the gases which come off will contain ammonia, which can be smelled. Test the gas also with moist red litmus paper.

121. **Phosphorus in seed.** Crush 25 kernels of wheat in a mortar. Place the material in an iron or porcelain crucible and heat it strongly over the Bunsen burner. Do not put the cover on. When cool, transfer the charred mass to a small beaker. Add 10 cc. of nitric acid and 50 cc. of water, and boil for 10 minutes. Break up the charred particles with a stirring rod during the boiling. Filter, and to half of the warm filtrate add 3 cc. of ammonium molybdate solution. A yellow precipitate shows the presence of phosphorus. Why?

122. **Calcium in clover.** Cut up some clover hay with a pair of shears. Place 25 grams of it in an iron crucible and ignite it over the burner till only ash is left. Transfer the ash to a beaker. Add 50 cc. of water and 5 cc. of nitric acid and then heat and finally filter. To about 20 cc. of the filtrate add ammonia water until the solution just turns red litmus blue. Then add 5 cc. of acetic acid, or more if necessary, until the reaction is distinctly acid to litmus. Now add 2 to 4 cc. of a saturated solution of ammonium oxalate. A white precipitate shows the presence of calcium in the ash. The precipitate is calcium oxalate, CaC_2O_4.

123. Potassium in plants. Test the ash from the clover for potassium with the flame test and blue glass as described in experiment 56.

124. Preparation of pectin. Grate a turnip and place the material in a linen bag and wash out all the soluble matter. Place the washed pulp in dilute hydrochloric acid (1 part of concentrated acid to 15 of water) and let stand 48 hours. Then squeeze out the acid liquid. Filter the liquid and to the filtrate add an equal bulk of alcohol. The precipitate formed is pectin. The presence of this substance in fruits is the cause of their jellying.

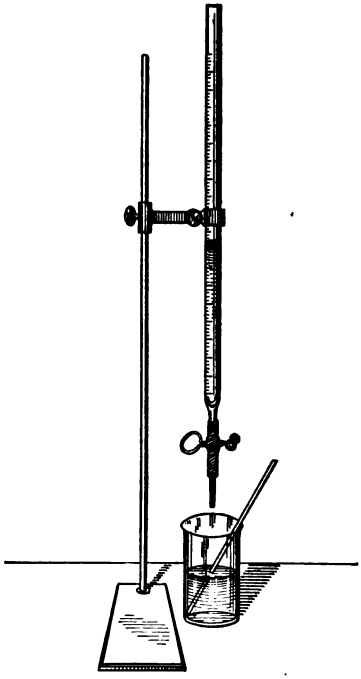


FIG. 127. — Neutralizing an acid.

125. Acids in foods. Vinegar contains acetic acid. Taste some vinegar, and then place 25 cc. of it in a beaker. Add two drops of phenolphthalein solution, and then carefully add a dilute solution of baking soda from a burette, Fig. 127, until a red color just appears. Now taste the liquid. It is no longer acid in taste. The soda has combined with the acid to form sodium acetate, a salt which has but little taste. What gas escaped? Baking soda is often used to neutralize excessive acidity in cookery, especially when acid fruits, such as apples, plums, or lemons, are employed.

EXPERIMENTS TO ACCOMPANY CHAPTER XVIII

126. Digestion of proteins. Get a hog's stomach at a slaughterhouse. Wash the stomach thoroughly with water. Remove the mucous membrane. Grind this membrane in a sausage grinder and place it in 500 cc. of 0.2 per cent solution of hydrochloric acid. Add 10 cc. of chloroform to prevent putrefaction. After standing at ordinary temperatures for 24 hours, filter the solution. Into 100 cc. of this solution place about 10 grams of coagulated egg white made by placing egg white into boiling water. Set the dish and its contents in a warm place, not above 100° F., for 24 hours. Examine the result. The protein has partly, if not entirely, disappeared. It has been digested by the pepsin from the stomach.

127. Action of malt on starch. In a mortar crush 30 malted barley kernels. Put this powder in a small flask, and add 15 cc. of water. After 24 hours filter off the solution, and to the liquid add 2 grams of flour and 100 cc. of water. Set the flask away in the cupboard for 24 hours. Then filter off the solution, and test a portion of the residue for starch with iodine solution. Malt converts starch to sugars. Test some of the solution for sugar with Fehling's solution as in experiment 54. If brewer's malt cannot be had, the barley grains themselves may be germinated; and when the sprouts are one-half to one inch long the seeds can be used in this experiment.

128. Action of saliva on starch. To obtain the saliva for this experiment, chew a small piece of pure paraffin, and collect the saliva in a beaker. Make some starch paste by suspending about 10 grams of starch in cold water and then stirring in boiling water. Take 25 cc. of this starch paste and adjust its temperature to about 104° F. Add 5 drops of

saliva and stir thoroughly. At intervals of a minute, remove a drop of the solution to a white tile and test the drop by means of a drop of iodine solution. If the blue color with iodine still forms after five minutes, add another 5 drops of saliva and stir. The turbidity of the starch solution should soon disappear, indicating the formation of soluble compounds which give no blue color with iodine. After further action, test the solution with Fehling's solution. A reddish brown precipitate with this reagent shows that the starch has been converted to the sugar, maltose.

EXPERIMENTS TO ACCOMPANY CHAPTER XIX

129. Water in foods. Mark two porcelain crucibles with your initials, using a lead pencil. Place them in the flame and heat them to redness for five minutes. Cool them in a desiccator and weigh them. Into each crucible place 4 grams of flour. Place them in an oven at 95° to 100° C. and allow them to remain for five hours. Cool them in the desiccator for one-half hour, and weigh them. The loss in weight, divided by the weight of the sample (4 grams) and multiplied by 100 gives the per cent of water in the flour. Flour will contain from 10 to 15 per cent water. Repeat the experiment with grated turnips, or with butter. Instead of porcelain crucibles ordinary tinned iron salve box covers may also be used.

130. Starch and protein in bread. Make the iodine (starch) and protein tests on samples of bread. Take some of the brown crust of the bread and grind it up and extract it with water. Test this extract with Fehling's solution. If you get a test for sugar, make a similar test, using an extract made from a part of the interior of the same loaf of bread. Compare the results. What are your conclusions?

131. **Effect of heat on potato starch.** With the point of a knife scrape lightly the surface of a raw potato and place a drop of the starchy juice upon a microscope slide. Cover this liquid with a cover glass and examine it under the compound microscope. Draw the starch grains. Make this examination, using a boiled potato and also a baked potato, drawing the starch grains in each case. Baking and cooking rupture the cells and thus it becomes easier to digest the food they contain.

132. **Starch grains.** With a compound microscope examine the starch grains of wheat, corn, rice, and oats. Draw the grains and compare these drawings with those of the different starch grains pictured in Figs. 86, 87, and 88. Proceed as follows: Place a drop or two of the starch water made from the grain on a slide. Cover the liquid with a cover glass, and examine the material under the microscope.

133. **Breakfast foods.** Under the microscope, examine two samples of cereal breakfast foods. Compare what you see with the appearance of the starch grains of wheat, corn, and oats, experiment 132. Tell from what grains the breakfast foods were made.

134. **Eggs. Composition of eggshell.** Examine a portion of the shell under the microscope, and note its physical character. Crush and grind an eggshell. Extract it thoroughly with warm water, and then dissolve the extracted mass in dilute hydrochloric acid. Observe that a gas comes off. Hold in the gas a drop of limewater on the end of a glass rod and note how cloudy the liquid becomes. Filter the acid solution, add ammonia water until strongly alkaline, and then add 5 cc. of ammonium oxalate solution. What is the precipitate? What is the gas that was formed? Why must chickens be fed limestone? Try the same experiment

with a piece of limestone; and also with wood ashes in place of eggshell.

135. **Egg albumen.** Prick two holes in an egg and blow out the white. Shake up a portion of it with water and note that it dissolves. Then boil some of the solution. The white coagulates. White of egg is a protein, and serves to nourish the growing chick during the incubation of the egg.

136. **Yeast.** On a watch glass mix thoroughly a piece of yeast of the size of a grain of wheat with about 5 cc. of water. Then with a stirring rod place a drop of this solution on a microscope slide, adding a drop of very dilute methyl violet solution. Cover with a cover glass, and examine under the microscope. The living, active cells appear colorless, while the decayed and lifeless ones are stained. Yeast cells are circular or oval in shape.

137. **Carbon dioxide from baking powder.** Baking powders are used for the production of carbon dioxide. Place 20 grams of the dried powder in a 250-cc. flask, provided with a cork through which passes a delivery tube, having its outer end below the surface of 100 cc. of limewater placed in a narrow beaker. To preserve the limewater from contact with the air, cover with one-fourth inch of kerosene. When water is added to the baking powder, carbon dioxide gas is rapidly evolved. The gas produces a precipitate in the limewater. What is the precipitate?

138. **Alum in baking powders.** Burn two grams of the baking powder in a porcelain dish. Extract the ash with boiling water and filter. Add to the filtrate enough ammonium chloride solution so that the mixture smells distinctly of ammonia. The appearance of a white, flocculent precipitate, especially upon warming, indicates the presence of alum in the powder. What is the precipitate?

139. **Testing baking powders for phosphates.** Dissolve

one-half gram of baking powder in 5 cc. of water and add 3 cc. of nitric water. Filter, and add 3 cc. of ammonium molybdate solution and warm gently. A yellow precipitate indicates the presence of phosphates.

140. **Fat in meat.** Grind up some meat in a meat grinder. Place 20 to 30 grams in a flask and cover it with 50 cc. of gasoline. Stopper the flask loosely. Shake it occasionally and let it stand one-half hour. Pour the gasoline off into a beaker and allow it to evaporate slowly. For this purpose it may be set outdoors. Observe the fat left behind in the beaker. Do the same experiment with peanuts instead of meat, grinding them up well before pouring on the gasoline. What are the results?

141. **Action of iron compounds on tannic acid.** Make an infusion of tea by placing 5 grams of tea in 100 cc. of boiling hot water and stirring well. Filter off some of the infusion, and add 5 cc. of ferrous sulphate solution (made by dissolving 1 gram of ferrous sulphate in 10 cc. of water and filtering). The solution turns dark in color; and if concentrated enough it becomes black. Try the same experiment with an infusion of hemlock or oak bark. This is the way common ink is made, only nutgalls are used instead of tea, or the bark of trees. All of these plant materials contain tannin.

EXPERIMENTS TO ACCOMPANY CHAPTER XX

142. **Microscopic examination of milk.** Place a drop of milk on a microscope slide. Cover with a cover glass and examine the milk for fat globules and for impurities, such as hair, dirt, etc. Make drawings. Examine a drop of skimmed milk. How do the two differ?

143. **Acidity of milk.** Test fresh milk with blue litmus paper. Let some of the milk sour and test this with blue

litmus paper. The acid formed is lactic acid, produced from the sugar in the milk by bacteria.

144. **Formation of lactic acid.** Place 5 grams of milk sugar in 100 cc. of water, add 5 cc. of skimmed milk and 2 or 3 crystals of sodium phosphate. Add 2 or 3 cc. of blue litmus solution. Leave the flask uncorked in the cupboard for 24 hours. Has the solution turned red? If it has, lactic acid (the acid in sour milk) has been formed. Taste the solution.

145. **Babcock test for fat in milk.** Measure with the pipette into the test bottle 17.6 cc. of milk. The sample should be carefully taken and well mixed. Add 17.5 cc. of commercial sulphuric acid (sp. gr. 1.83) from the cylinder; mix the acid and milk by rotating the bottle. Then place the test bottle in the centrifugal machine and whirl 5 minutes at a rate of 800 to 1200 revolutions per minute. Add sufficient hot water to the test bottle to bring the contents up to the shoulder, and whirl two minutes. Then fill the bottle with water to near the top of the graduations and whirl again for two minutes, and then read the fat. Read from the extreme lowest to the highest point. Each large division, as from one to two, represents a whole per cent. Each small division represents 0.2 of a per cent. Test a Holstein milk and a Jersey milk. Test the milk from the cows on your farm.

146. **Formalin in milk.** The instructor should place some of the preservative in milk and have the pupils detect it. To 10 cc. of milk add 10 cc. of a solution of concentrated hydrochloric acid containing per liter 2 cc. of 10 per cent ferric chloride solution. Pour the mixture in a teacup and heat it slowly to boiling over a flame, giving the cup a slight rotary motion. If formalin is present, a violet coloration will appear.

147. **Test for oleomargarine.** Foam test for purity of

butter. Heat about 3 grams of the sample in a large iron spoon over a low flame, stirring constantly with a splinter. Genuine butter will boil quietly, with the production of considerable froth or foam. Oleomargarine and renovated butter will sputter and make much more noise. Always compare by making a simultaneous test with genuine butter.

148. **Waterhouse test.** Into a small beaker pour 50 cc. of sweet milk, heat nearly to boiling, and add 5 to 10 grams of butter or oleomargarine. Stir with a glass rod until the fat is melted. Then place the beaker in cold water and stir the milk until the temperature falls sufficiently for the fat to congeal. At this point the fat, if the sample is oleomargarine, can be collected into one lump by means of a rod. If it is butter, it will granulate and cannot be thus collected.

EXPERIMENTS TO ACCOMPANY CHAPTER XXI

149. **Purity of Paris green.** Take about one gram of Paris green. Put it in a beaker and add 25 cc. of ammonia water. Stir and let it stand for five minutes. If the "green" is pure, it will form a clear, dark blue solution, and leave no solid residue.

150. **Preparation of lime-sulphur.** Slake 25 grams of lime in 200 cc. of water. Add 50 grams of flowers of sulphur. Shake the mixture vigorously, and heat it in a steam bath with frequent shaking. Finally place it over a wire gauze or a free flame and boil for half an hour. The mixture must be watched and occasionally stirred in order to avoid violent bumping. After boiling, allow all to stand and settle. Note the color of the liquid and the character of the sediment.

151. **Preparation of Bordeaux mixture.** Dissolve 0.5 pound of copper sulphate in 2 pounds of hot water. Also slake 0.3 to 0.4 pound of lime in 2 pounds of water and strain

the liquid through a cheesecloth into a pail so as to remove the coarse material. The solution of copper sulphate is then poured into the pail and all is well stirred. In the preparation of Bordeaux mixture it is the aim to use just enough lime to combine with all of the copper sulphate. Test the mixture with litmus paper. If it is acid, enough more lime should be added to turn the litmus blue. The ferrocyanide test can also be performed as follows: Filter some of the finished mixture, and add a few drops of potassium ferrocyanide. If a reddish brown precipitate or color appears, more lime should be added. What is the reddish brown precipitate?

152. **Destruction of germinating power by formalin.** Place about 6 or 8 kernels of different grains, such as corn, barley, and wheat, in a solution of formalin of 20 per cent strength. Leave them there for half an hour. Remove the seeds, wash with water, and place them in moist sawdust for germination. Do they germinate? What care must be practiced in treating seeds with such poisons?

LISTS OF APPARATUS AND CHEMICALS NECESSARY FOR THE EXPERIMENTS DESCRIBED IN THIS CHAPTER

General Laboratory Apparatus

1 Babcock tester.
1 Microscope.
1 Inexpensive balance sensitive to 1 cg., and set of weights, 50 gr. to 1 cg.
1 Drying oven (copper).
1 Water bath (6 holes).
2 Mortars (4 in.), preferably of porcelain, though glass mortars will do.
2 Thermometers (0°–120°).
4 Burettes, 50 cc., and holders.
2 Condensers (Liebig).

1 Gross corks, sizes 7, 8, 9, 10, and 12.
1 Cork borer, set of six.
2 500-cc. cylinders.
6 Cobalt glasses 10 cm. square.
1 Magnet (horseshoe).
4 Pneumatic troughs, galvanized iron, 1 ft. long, 8 in. high, and 8 in. wide, or can use 1-gal. stoneware milk pans.
3 Platinum wires, 4 in. long, about No. 26 B. & S. guage.
1 Blast lamp, bellows and rubber tubes.
10 lb. Glass tubing, sizes up to $\frac{1}{2}$ in. int. diam.
4 Retorts (glass) 500 cc.
1 Dozen cylinders 100 cc.

Tacks, nails, brads, carpenters' tools, etc., which can always be obtained from local dealers, have not been mentioned here. With the aid of carpenters' and tinners' tools many useful pieces of apparatus can be constructed by the students. Often the local tinner will make a piece of apparatus that will serve as well as one secured elsewhere at higher cost.

List of Apparatus used by Each Student

1 Bunsen burner and tubing, or alcohol lamp.
2 Stirring rods.
4 Watch glasses (diam. 3 in.).
2 Erlenmeyer flasks (300 cc.).
25 Filter papers (11 cm.).
1 Box matches.
1 Wire gauze (4 in. square).
2 Crucibles (porcelain, diam. 3.5 cm.)
1 Test tube brush.
2 Funnels (diam. 6.5 cm.).
1 Sand bath (4 in. diam.).
1 Test tube stand.
1 Wooden stand for funnels.
12 Test tubes 12 cm. long, diam. about 1.7 cm.
6 Beakers (nest 100 to 600 cc.).
1 Casserole (200 cc.).
2 Evaporating dishes (diam. 7.5 cm.).
1 Crucible tongs (9 inches).

1 Test tube clamp.
1 Water bath (copper).
2 Aluminum dishes (2½ in. diam.).
2 Feet of glass tubing (soft), ext. diam. 6 mm.
1 File (15 cm. long), triangular.
1 100-cc. cylinder.
1 Bottle each of blue and red litmus paper.
1 Ring stand and rings.
1 Deflagrating spoon.
4 Bottles (wide mouth) 250 cc.
Rubber tubing 2 ft. (int. diam. 5 mm.).
2 Flasks 500 cc.
1 Thistle tube.
1 Desiccator.
1 Iron crucible 50 cc.

Chemicals for Ten Students

	Grams
Acid, Hydrochloric, sp. gr. 1.2	1500
Nitric, sp. gr. 1.4	1000
Sulphuric, sp. gr. 1.84	2000
Acetic (50%)	200
Alcohol	100
Alum (potassium)	25
Aluminum chloride	25
Ammonium carbonate	50
Ammonium chloride	250
Ammonium hydroxide, sp. gr. 0.9	1500
Ammonium oxalate	25
Ammonium nitrate	100
Ammonium sulphate	100
Ammonium molybdate	50
Arsenic trioxide	100
Barium chloride	20
Barium hydrate	20
Bone black	100
Bone ash	100
Borax	50
Calcium chloride (granular)	200

	Grams
Calcium sulphate (gypsum)	50
Calcium carbonate (marble)	50
Carbon bisulphide	30
Citric acid	25
Charcoal, ten pieces 8 cm. × 4 cm.	
Cochineal	5
Cobalt nitrate	10
Copper (turnings or scrap)	250
Copper oxide (black)	50
Copper sulphate	100
Formalin	50
Fuchsine	25
Glucose (sirup)	1000
Iron chloride	30
Iron powder	25
Iron sulphate (ferrous)	75
Iron sulphide	250
Iodine	10
Indigo	15
Lime	50
Logwood	10
Lead acetate	25
Lead (cuttings from lead pipe)	100
Lead nitrate	10
Litmus	10
Magnesium carbonate	50
Magnesium sulphate	20
Magnesium wire or ribbon	10
Manganese dioxide	1000
Methyl violet	5
Mercury	100
Mercuric chloride	20
Paraffin	10
Phenolphthalein	5
Phosphorus, yellow	10
Potassium carbonate	50
Potassium chlorate	250
Potassium sulphate	20
Potassium dichromate	25

	Grams
Potassium hydroxide	50
Potassium iodide	50
Potassium nitrate	25
Potassium permanganate	10
Potassium ferrocyanide	25
Rochelle salts	20
Silver nitrate	10
Sodium (metallic)	5
Sodium bicarbonate	20
Sodium carbonate	100
Soda-lime	100
Sodium hydrogen phosphate	20
Sodium hydroxide	400
Sodium nitrate	200
Sodium silicate (water glass)	100
Sodium sulphate	10
Strontium chloride	10
Sulphur	200
Tannic acid	20
Tin	30
Zinc (granulated)	1000
Zinc sulphate (crystals)	100

This list does not include common materials that are always easily obtained, as sugar, salt, lard, tallow, clay, starch, cotton, iron, wire, gasoline, benzine, chloroform, bleaching powder (chloride of lime), tartar emetic, turpentine, etc.

For the convenience of teachers the names of a number of firms from whom apparatus and chemicals may be purchased are here given: E. H. Sargent & Co., Chicago, Ill.; Eimer and Amend, New York City; Bausch and Lomb, Rochester, N. Y.; Arthur H. Thomas Co., Philadelphia, Pa.; Scientific Materials Co., Pittsburg, Pa.; Henry Heil & Co., St. Louis, Mo.; Denver Fire Clay Co., Denver, Col.; Kny-Scheerer Co., New York City.

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